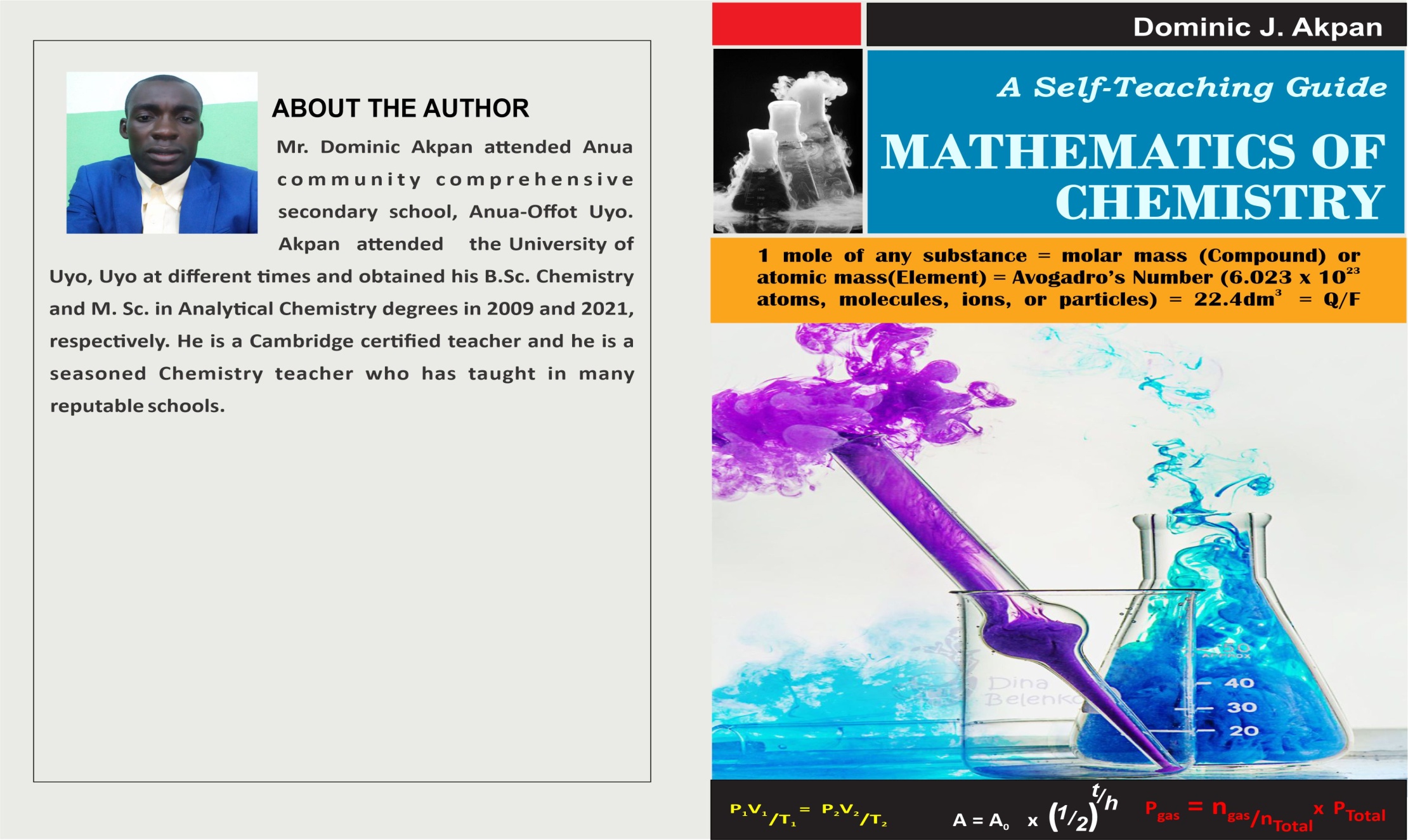
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**ABOUT THIS BOOK**

* It is written to boost students' learning ability in chemistry calculations and make learning of Chemistry fun.
* It is written by an experienced teacher and examiner.
* The book contains all necessary Chemistry calculations for Senior Secondary Schools to support the attainment of the highest grade in Chemistry.
* It provides all the mathematical explanations you need to calculate efficiently in Chemistry examinations.
* It contains more solved examples for students to master the topics covered. Take away questions (with immediate answers) are given to allow the students test their knowledge in the subject.
* This book is designed to assist students in gaining a commendable grade in Chemistry.

**HOW TO TEACH CALCULATIONS IN CHEMISTRY (For teachers)**

**A few steps in teaching calculations in Chemistry:**

1. Teaching calculations takes time. It should not be rushed.

2. Because of the way we learn things, teaching calculations is best done using a stepwise, gradual approach. Rather than spend a week on calculations, and then not revisit the topic until just before the examination, it is more productive to introduce a calculation topic, give some practice in the process and then set time aside to return to it from time to time over the next few weeks. Learning often occurs at second or third encounter because, at first encounter the mind is too busy trying to assemble the new information and make sense of it all.

3. The method known as see one => try one => do one should be applied when teaching Chemistry calculations. That is to say, the teacher explain, with the usual questioning, how to carry out a new type of calculation. The class then attempts a calculation with a close support from the teacher. Once the teacher is satisfied that most students are reasonable confident, further calculations can be attempted with support being provided where necessary.

4. The difficulty of a calculation type should advance by small steps when teaching it. To take an example, supposed that the reacting masses calculation is being introduced. In explaining the method, it is advisable to use a simple balanced chemical equation as possible, like Fe + S => FeS. After explaining how to do the calculation, students should attempt one. The equation chosen should be something similar, say Zn + S => ZnS.

5. Practice daily. One of the key areas in the study of Chemistry is practice. Solving calculations should be a feature part of your study routine.

**IMPORTANT FORMULAS IN CHEMISTRY**

**11.**

**19.**

**21. Energy = n × ΔH**

**22. ∆G = ∆H - T∆S**

**23. ∆S =**

**25. AF = Aoe-kt**

CHAPTER ONE

CHEMICAL FORMULAE AND EQUATIONS

**INTRODUCTION**

Chemistry calculations are based on chemical formulas and equations. Chemists use algebraic equations to express relationships between quantities. An example of such an equation is the relationship between the density, mass, and volume of a substance: (density is equaled to mass divided by volume). Given (or making) measurements of the mass and volume of a substance, you can use this equation to determine the density. The substances undergoing reaction are called **reactants**, and their formulas are placed on the left side of the equation. The substances generated by the reaction are called **products**, and their formulas are placed on the right side of the equation.

Plus signs (+) separate individual reactant and product, formulas, and an arrow (→) separates the reactant and product (left and right) sides of the equation. The relative numbers of reactant and product species are represented by coefficients (numbers placed immediately to the left of each formula). A coefficient of 1 is typically omitted.

**VALENCY AND CHEMICALFORMULA**

Valency is the combining power or capacity of an element or radical. Valencies are whole numbers representing the numbers of hydrogen atoms required to combine with an atom for it to attain the stable electronic configuration. Elements having a single valency are said to be monovalent, while those with two valencies are said to be divalent. Elements with more than two valencies are said to be multivalent. Copper, for example, is a divalent element with valencies of 1 and 2.This accounts for the reason why it can form two different compounds with the same element. Examples are copper (I) oxide and copper (II) oxide. The valencies of some substances are given in the Table below:

|  |  |  |  |
| --- | --- | --- | --- |
| **Substance** | **Valency** | **Substance** | **Valency** |
| Al | 3 | Group 2 | 2 |
| H | 1 | OH- (hydroxide or hydroxyl ion) | 1 |
| Pb | 2 and 4 | CO32- (trioxocarbonate (iv) ion) | 2 |
| Ag | 1 | NO3-  (trioxonitrate (v) ion) | 1 |
| Cl | 1 | PO43- (tetraoxophosphate (v) ion) | 3 |
| Zn | 2 | SO42- (tetraoxosulphate (vi) ion) | 2 |
| S | 2 and 6 | HCO3- (hydrogen trioxocarbonate iv) | 1 |
| O | 2 and 6 | NH4+ (ammonium ion) | 1 |
| Group 1 | 1 | Cr2O72- (heptaoxodichromate vi) | 2 |

**Table 1: Valencies of some common substances**

**Example1:** Deduce the chemical formula of hydrogen chloride.

**Solution**

The constituent elements are hydrogen and chlorine, both of which have a valency of 1. Thus, we should write H1Cl1. Since the valencies are equal, then we omit them and write the chemical formula of hydrogen chloride as HCl.

H CL

1 1

**Example 2:** Deduce the chemical formula of tetraoxosulphate (VI) acid.

**Solution**

The acid contains hydrogen and tetraoxosulphate (VI) (a radical). The valency of hydrogen is 1 and that of tetraoxosulphate (VI) is 2.

H SO42-

1 2

Since the valencies are not equal, then we interchange the valencies and use them as subscripts to obtain H2SO4.

Note that SO4 is not enclosed in parentheses because it has a subscript of 1.

**Example 3:** Deduce the chemical formula of iron (III) tetraoxosulphate (VI).

**Solution**

The constituent element and radical are iron and tetraoxosulphate (VI) respectively.

Fe SO42-

3 2

Since the valencies are not equal, then we interchange the valencies and use them as subscripts to obtain Fe2(SO4)3. The Roman numeral of 3 appearing in the name of the compound shows that the valency of iron is 3.

Note that SO4 is enclosed in parentheses because its subscript is greater than 1.

**It is common practice to use the smallest possible whole-number coefficients in a chemical equation.**

**The chemical formula of a compound can easily be deduced from the valencies of its constituent elements by adhering to the following rules:**

•Write the symbols of the constituent elements and radicals of the compound, with the valency of each denoted as its superscript to the right. The charges of radicals should be ignored. **A radical (or polyatomic ion) is a group of atoms that reacts as a unit.**

•If the valencies are not equal, they should be balanced by interchanging the superscripts and using them as the subscripts of the right of the chemical symbols.

•If a radical is involved, then it must be treated as a unit by enclosing it in parentheses before the subscript is added.

This is a requirement the equation must satisfy to be consistent with the law of conservation of matter.

**Types of Chemical Equations**

**A word equation** uses words to represent the reactants and products in a chemical reaction .E.g Zinc metal + hydrogen tetraoxosulphate (VI) acid ==> Zinc tetraoxosulphate (VI) + hydrogen gas

**A molecular equation** is an equation in which the formulas of the compounds are written as though all substances exist as molecules. However, there is a better way to show what is happening in this reaction. All of the aqueous compounds should be written as ions. A formula equation or molecular equation uses chemical symbols or formulas, but does not reveal the ratios of the products and reactants. E.g

Zn + H2SO4 ==> ZnSO4 + H2

**An ionic equation** is an equation in which dissolved ionic compounds are shown as free ions. A balanced ionic equation shows the reacting ions in a chemical reaction.

Na+ + Cl- + Ag+ + NO3- ==>Na+ + NO3- + AgCl

**A spectator ion** is an ion that does not take part in the chemical reaction and is found in solution both before and after the reaction. In the above reaction, the sodium ion and the nitrate ion are both spectator ions and as such they are cancelled out.

**TAKE AWAY**

1. Deduce the chemical formula of potassium permanganate. **Ans. KMnO4**

2. The chemical formula of aluminium oxide. **Ans. Al2O3**

3. Deduce the chemical formula of calcium hydroxide. **Ans. Ca(OH)2**

4. Deduce the chemical formula of tin (II) chloride. **Ans. SnCl2**

5. Deduce the chemical formula of lead (II) nitrate (v). **Ans. Pb(NO3)2**

**CHAPTER TWO**

**RELATIVE ATOMIC AND MOLECULAR MASSES**

**RELATIVE ATOMIC MASS**

The masses assigned for elements in the periodic table are relative masses in terms of atomic mass units (amu). The atomic mass unit is based on a relative scale in which the reference is the carbon-12 isotopes which is assigned a mass of exactly 12 amu.

**Definition:** *The relative atomic mass of an element is a number which shows how heavy an atom of that element is compared with an atom of another element. The modern method of finding relative atomic mass (R.A.M.) is to use an instrument called a* ***mass spectrometer***.

Relative Atomic Mass (R.A.M)

***Difference between mass and weight***

It is very important that we understand the difference between mass and weight. ***Mass is an invariant measure of the amount of matter in an object while weight is the forces of attraction between an object and its surrounding.*** The mass of a substance remain constant regardless of where you measure it while the weight of the substance may change. R.A.M have no unit since they are ratios of two masses. Relative atomic masses of some common elements in the periodic table are given below:

H = 1.0, Li = 7.0, = 9.0, B = 11.0, C = 12.0, N = 14.0, O = 16.0, F = 19.0, Na = 23.0, Mg = 24.0, Al = 27.0, Si = 28.0, P = 31.0, S = 32.0, Cl = 35.5, K = 39.0, Ca = 40.0, Cr = 52.0, Mn = 55.0, Fe = 56.0, cu = 64.0, Zn = 65.0, Ag = 108.0, Pb = 207.0 etc.

***Relative Molecular Masses***

*A molecule consists of a combination of two or more atoms of the same kind (H2, N2, O2) or different kinds of atoms (NH3, HCl, H2O etc).The molecular mass of a compound is the sum of the atomic masses of all the atoms in the correctly written formula.* Molar mass (molecular mass) is the mass in grams that contains 6.023 × 1023 atoms of the element. The molar mass of a substance = 1 mole of the substance (6.023 × 1023 atoms/ions/particles). It has the unit ***gmol***-1

* 1 mole of carbon = 12g = 6.023 × 1023 atoms of carbon.

***Note*: *Relative molecular mass is calculated same as molecular mass but there is no unit attached to it. It has no unit since it is only a ratio.***

If the individual atomic masses of all the atoms in a formula are added together, you have calculated the relative molecular mass.

Example, for ionic compounds e.g. NaCl. The relative molecular mass of NaCl = 23 + 35.5 = 58.5) or molecular mass for covalent elements or compounds e.g. Relative Molecular mass of N2 = 28 from (2 x 14) or compounds e.g. Relative molecular mass of glucose (C6H12O6) = 180 from [(6 x 12) + (12 x 1) + (6 x 16)]

In a balanced chemical symbol equation, the total of relative Molecular masses of the reactants is equal to the total relative Molecular masses of the products (law of conservation of mass).

The shorthand **R*MM*** can be used for the formula of any element or compound.

RMM = relative molecular mass = the sum of all the atomic masses for all the atoms in a given formula.

**NOTE:** You cannot successfully calculate formula/molecular masses if you cannot read a formula correctly. Whereas relative atomic mass only applies to a single atom, anything with at least two atoms in the formula requires the term ***relative formula mass or relative molecular mass*** to be used.

**WARNING: The most common error is to use atomic/proton numbers instead of atomic masses, unfortunately, except for hydrogen (where relative atomic mass is the same as atomic/proton number) they are different.**

***Solved example 1:***

Calculate the relative molecular mass of carbon dioxide.

***Solution***

The correct formula: CO2

1- atom of carbon + 2 - atoms of oxygen

1 × 12 + 2 × 16 = 44

***Solved example 2:***

What is the molecular mass of calcium tetraoxosulphate (VI)?

***Solution***

The correct formula: CaSO4

1-atom of calcium + 1-atom of sulphur + 4-atoms of oxygen

1 × 40 +1 × 32 + 4 × 16 = 136.0

***Solved example 3***

What is the relative molecular mass of sodium chloride?

***Solution***

The correct formula: NaCl

1 - atom of sodium + 1 - atom of chlorine

1 × 23 + 1 × 35.5 = 58.5

***Solved example 4***

Workout the molecular masses of the compounds below:

1. Ca(OH)2
2. NaOH
3. Mg(NO3)2

***Solution***

1. Calcium hydroxide/lime water ; Ca(OH)2

1-atom of calcium + 2- atoms of oxygen + 2 - atoms of hydrogen

1 × 40 + 2 ×16 + 2 × 1 = 74.0

1. Soda ash/sodium hydroxide; NaOH

1-atom of sodium + 1- atom of oxygen + 1- atom of hydrogen

1× 23 + 1 ×16 + 1 ×1 = 40

1. Magnesium trioxonitrate (V); Mg(NO3)2

1-atom of magnesium + 2 - atoms of nitrogen + 6 - atoms of oxygen

1 × 24 + 2 × 14 + 6 × 16 = 148

**Solved example 5:** Calculate the molar mass of sodium trioxocarbonate (IV) decahydrate.

**Solution**

The formula for sodium trioxocarbonate (IV) decahydrate is Na2CO3.10H2O.

2 - atoms of sodium + 1 - atom of carbon + 3 - atoms of oxygen + 10 - molecules of water(H2O)

Molar mass of Na2CO3.10H2O = 2 × 23 + 1 × 12 + 3 × 16 + 10 (2 + 16) = 286 gmol-1.

**ISOTOPES**

Isotopes are atoms whose nuclei have the same atomic number but different mass numbers i.e the nuclei have the same proton but different neutrons. Examples of elements that exhibit isotopy (a phenomenon in which atoms have the same proton number but different neutron numbers) are carbon, oxygen, hydrogen, chlorine, neon, nitrogen etc.

***Note:* The relative atomic masses of elements which exhibit isotopy are usually not whole numbers because of the relative abundance of the isotopes.**

***Isotopic calculations***

Given the percentage abundance and the mass numbers, one can easily obtain the relative atomic mass of elements using the formula below:

Relative Atomic Mass (R.A.M.) =

Where %1, %2 and %3 are percentages of the first, second and third isotopes respectively. M1, M2 and M3 are masses of the first, second and third isotopes respectively.

***Solved example 1***

Calculate the relative atomic mass of an element whose isotopes have relative abundance as follows:

28Y:92.2%, 29Y: 4.7%, 30Y:3.1%.

***Solution***

Using the formula R.A.M. =

R.A.M. =

R.A.M. =

R.A.M. =

R.A.M. = 28.1(it is not a whole number because of the presence of three isotopes of Y)

***Solved example 2***

If an element Z has two isotopes of mass numbers 18 and 20 respectively and these are present in the proportion of 1 : 2;find the relative atomic mass of Z.

***Solution***

Total ratio = 1 + 2 = 3

R.A.M.=

R.A.M. =

R.A.M. =

R.A.M. =

R.A.M. =3 33

***Solved example 3***

Chromium, Cr has the following isotopic masses and fractional abundances:

**Mass number isotopic mass (amu) Fractional abundance**

50 49.9461 0.0435

52 51.9405 0.8379

53 52.9407 0.0950

54 53.9389 0.0236

What is the atomic mass of chromium?

***Solution***

Multiply each isotopic mass by its fractional abundance, then sum:

49.9461 amu × 0.0435 = 2.17 amu

51.9405 amu × 0.8379 = 43.52 amu

52.9407 amu × 0.0950 = 5.03 amu

53.9383 amu × 0.0236 =1.27 amu

51.99 amu

Hence, the atomic mass of Cr = 51.99 amu

**Example 4:** Bromine consists of two isotopes, 50% 79Br and 50% 81 Br, calculate the relative mass of bromine from the mass numbers.

**Solution**

Think of the calculation in terms of 100 atoms



So the relative atomic mass of bromine is 80 or R.A.M (Br) = 80

**Example 5:** Chlorine consists of two isotopes, 75% chlorine - 35 and 25% chlorine - 37.

**Solution**

Again think of the data based on 100 atoms, so 75 have a mass of 35 and 25 atoms have a mass of 37.



So the relative atomic mass of chlorine is 35.5 or R.A.M (Cl) = 35.5

**Example 6:** Europium atoms consist of 47.8% Eu-151 and 52.2% of Eu-153. Calculate the relative atomic mass of europium.

**Solution**

152.0

The relative atomic mass of europium is 152.0 (to 1 decimal place).

**Example 7:** Atoms of the element silicon consist of 92.2% silicon- 28, 4.7% silicon-29 and3.1%ofsilicon-30. Calculate the relative atomic mass of silicon.

**Solution**

****

The relative atomic mass of silicon is 28.1 (to 1 decimal place or 3 significant figures).

***It is possible to do the reverse of a relative atomic mass calculation if you know the relative atomic mass and which isotopes are present. It involves a little bit of arithmetical algebra.***

**Example8:** The relative atomic mass (RAM) of boron is 10.81 and consists of only two isotopes, boron 10 and boron-11. Calculate the percentages of the two isotopes.

**Solution**

The relative atomic mass of boron was obtained accurately in the past from chemical analysis of reacting masses but now mass spectrometers can sort out all of the isotopes present and their relative abundance.

% of the first isotope + % of the second isotope = 100. If you let X = % of boron 10, then100 – X is equal to % of boron-11.

R.A.M. =



100 ×10.81 =× 10 + [(100 - ) × 11]

1081 = 10+1100 - 11

1081 – 1100 = 10- 11

- 19 = -  (change sides change sign)

= 19

Recall, 100 – = % of boron-11.

% of boron-11 = 100 – 

% of boron-11 = 100 – 19

% of boron-11 = 81%

Therefore,= 19, so naturally occurring boron consists of 19% 10B and 81% 11B.

**TAKE AWAY**

* + - 1. Work out the relative molecular mass of these compounds:
         1. Sulphur (iv) oxide. **Ans.64**
  1. Potassium trioxonitrate (v). **Ans. 101**
  2. Magnesium tetraoxosulphate (vi) heptahydrate. **Ans. 246**
  3. Aluminium tetraoxosulphate (vi). **Ans. 342 2.**

2. Calculate the relative formula mass of propanol, C3H8O, CH3CH2CH2OH (the same formula can be expressed in different ways). Relative atomic masses: C=12,H=1,O=16. **Ans. 60**

3. What is the relative formula mass of magnesium nitrate, Mg(NO3)2. Relative atomic masses: Mg = 24,N=14,O=16. **Ans. 148**

4. Calculate the relative formula mass of blue hydrated copper (II) teraoxosulphate(VI) pentahydrate crystals,CuSO4.5H2O.[Cu = 63.5,S = 32,H = 1,O = 16] **Ans. 159.5(160 if you use Cu = 64).**

5. Chlorine consists of the following isotopes:

Isotope isotopic mass (amu) fractional abundance

Cl – 35 34.96885 0.75771

Cl – 37 36.96590 0.24229

What is the atomic mass of chlorine? **Ans. 35.5**

1. An element X has two isotopes,and, present in the ratio 1:3. What would be the relative atomic mass of X? **Ans. 21.5**
2. Calculate the relative atomic mass of copper from its isotopic composition (isotope abundance). Naturally occurring copper consists of 69.2% copper – 63 (63Cu) and 30.8% copper - 65(65Cu). **Ans. 63.6**
3. Silver atoms consist of 51.4 % of the isotope 107Ag and 48.6% of the isotope 109Ag. Calculate the relative atomic mass of silver. **Ans. 108.0** (to 1 decimal place).
4. An element, Q, contains 69% of 63Q and 31% of 65Q. What is the relative atomic mass of Q? **Ans. 63.6**
5. The isotopes of neon are represented by the symbols 20 XNe, 21y Ne and 22z Ne. What is the relationship between X, y and z?  **Ans. X = y = z**

**CHAPTER 3**

**PERCENTAGE COMPOSITION BY MASS**

In finding the actual proportion by weight or mass (%) of an element in a compound, the weight of the part must be divided by the weight of the whole compound. For instance, a bolt and nut can be compared to a molecule with two atoms. If the nut weighed 100g and the boot weighed 200g, the whole would weighed 100 + 200 = 300g. The nut is a part of the whole. To find the percentage by weight or mass that the nut occupies with the whole, divide the weight of the nut by the weight of the whole and multiply by 100.

Hence, 

The percentage of the weight contributed by the nut is 33.3% of the weight of the whole.

***Percent composition is the proportion of an element present in a compound, found by dividing the mass of the element by the mass of the whole compound and multiplying by 100%.***

**The percentage by mass composition of a compound in terms of its constituent elements is calculated in three easy steps:**

1. Calculate the formula or molecular mass of the compound.
2. Calculate the mass of the specified element (for its %) in the compound, taking in to account the number of atoms of the element in the compound formula.
3. Calculate (ii) as a percentage of (i): relative atomic mass of element x number of atoms of element in formula × 100 divided by formula or molecular mass of the compound. % by mass of Z = 100 x Ar(Z) x atoms of Z/Mr(compound).

It always seems complicated when stated in this form, but the calculations are actually quite easy as long as you can correctly read a formula. Having known the formula of a compound, relative atomic mass and its molecular mass, one can work out the percentage by mass of each element present in the compound.

*% by mass* =

***Solved example 1:***

Calculate the percentage by mass of silicon and oxygen in silicon (IV) oxide (silica). [O =16, Si = 28]

***Solution***

***Hint***: Work out the relative molecular mass of silicon (IV) oxide.

The correct formula: SiO2

Relative molecular mass (R.M.M) of silicon (IV) oxide = 1-atom of silicon + 2-atoms of oxygen

R.M.M.= 1 × 282 × 16

= 28 + 32

= 60

*% by mass* =

*% of silicon in SiO2* =

*% of silicon in SiO2* =

*% of silicon in SiO2* = 46.7%

*% of oxygen in SiO 2* =

=

*% of oxygen in SiO2* = 53.3%

**Note:** 46.7% of silicon in SiO2 obtained from sand, quartz etc. This is an example of the law of constant composition, which states that all pure samples of a compound contain the same elements chemically combined in the same proportion by mass.

***Solved example 2***

Find the % by mass of magnesium, oxygen and sulphur in magnesium tetraoxosulphate (VI).

***Solution***

The correct formula: MgSO4

Relative molecular mass of MgSO4 (R.M.M)= 1 - atom of magnesium + 1-atom of sulphur + 4 - atoms of oxygen

Relative molecular mass of MgSO4 = 1× 24 + 1 × 32 + 4× 16

R.M.M = 120

*% of Mg* =

*% of Mg* =

*% of Mg* = 20%

*% of S* =

*% of S* =

*% of S* = 26.7%

*% of O* =

*% of O* =

*% of O* = 53.3%

***Solved example 3***

A compound was analyzed in the laboratory and found to contain 69.94% iron and 30.06% oxygen. Is the compound iron (II) oxide, FeO or iron (III) oxide, Fe2O3? (Fe =55.85, O = 16.0)

***Solution***

Molecular mass of FeO = 1-atom of Fe + 1-atom of O

Molecular mass of FeO = 1 × 55.85 + 1 × 16.0

Molecular mass of FeO = 71.85g

*% of Fe =* =

= /mol

*% of Fe =* =

***Solved example 4***

Calculate the percentage composition of each of the following:

1. C3H6
2. C5H10

***Solution***

1. Relative molecular mass of C3H6 = 312 + 61 = 42.0

H

= 1×6

= 14.3%

.

1. Relative molecular mass of C5H10 = 5 12 + 10 1 = 70.0

=

100

%

*The percentages are the same as those in part (a). This is so because the ratio of atoms of carbon to atoms of hydrogen is the same i.e 1:2 in both compounds.*

***Solved example 5***

A forensic scientist analyzes a drug and finds that it contains 80.25% carbon and 9.55% hydrogen. Could the drug be pure tetrahydrocannabinol (C21H30O2)?

***Solution***

Relative molecular mass of (C12H30O2) = 12 21 + 130 + 2 16 = 314.0

***Solved Example 6***

In an experiment 6.0g of metal **M** was burned in a crucible, by heating in air, until there was no more gain in weight. Apart from the mass of the crucible, the final mass of the residue was 10.0g. The oxide **O** formed was an essential ingredient in a ceramic pigment mixture **P** for glazing pottery.

(a)What % of the metal **M** inthe oxide **O**

(b)if the mixture **P** must contain 25% by mass of the metal **M**.

What mass of the oxide **O** is needed to make 12g of the mixture **P**?

**Solution**

1. What is % of the metal **M** inthe oxide **O**

= 60% of M in the oxide O

(b) 25% of 12g is 12 x 25/100 = 3.0g, so this is the mass of metal **M** in 12g of mixture **P** as the oxide **O**

You now have to scale up to the mass of the oxide **O** needed Which contains 60% of **M**.

Scaling up gives 3.0 x 100/60 = 5.0g of oxide **M** is needed.

***Solved Example 7:***

Calculate the percentage of water of crystallization in magnesium tetraoxosulphate (VI) crystals, MgSO4.7H2O, known as Epsom salt.

**Solution**

Molecular mass of Epsom salt = 24 + 32 + 64 + (7 x 18) = 246g/mol

Molecular mass of water = 18 (1 × 2 + 16), mass of seven water molecules is 7 x 18 = 126g/mol

Therefore, % of water of crystallization in the crystals = 100 x 126/246 = 51.2% H2O

***Solved Example 8***

Calculate the mass percentage of benzene and carbon tetrachloride, if 22g of benzene is dissolved in 122g of carbon tetrachloride.

**Solution**

We have masses of benzene (C6H6) and carbon tetrachloride (CCl4). So calculate the mass of the Solution by adding these two. Then calculate the mass percentage from the formula:

Mass of benzene = 22g; mass of CCl4 = 122g

Mass of solution = 22 + 122 = 144g

Mass % of CCl4 = 100 - 15.28 = 84.72%

OR

**Solved Example 9:** What is the percentage composition of carbon in two moles of calcium hydrogen trioxocarbonate (IV)?

**Solution**

Two moles of calcium hydrogen trioxocarbonate (IV) means 2 CaHCO3.

The molar mass of 2 CaHCO3 = 2[40 + 1 + 12 + 16 × 3] = 2[101] = 202gmol-1.

**Solved Example 10:** 50 tons ofimpure cast iron from the blast furnace contain 47 tons of iron. Calculate the percentage of the impurities in the cast iron.

**Solution**

***TAKE AWAY***

* + - 1. Find the percentage by mass of 0.368g in a limestone sample weighing 1.267g. **Ans. 29.08 %.**

2. Calculate the % by mass of elements in the following compounds

1. Calcium trioxocarbonate (IV). **Ans. Ca = 40%, C = 12%%, O = 48%**
2. Potassium hydrogen trioxocarbonate (IV). **Ans. K = 39%, H = 1%, C = 12%, O = 48%**
3. Ethanol. **Ans. C = 52.2%, H = 13.0%, O = 34.8%**
4. Aluminium sulphide. **Ans. Al = 36%, S = 64%**

e) Ethanoic acid. **Ans. C = 40%, H = 6.67%, O = 53.33%**

3. Ammonium tetraoxosulphate (VI), (NH4)2SO4, is an important ingredient in many artificial fertilizers.

(a) Calculate the percentage of nitrogen and the percentage of sulfur in ammonium tetraoxosulphate (VI). **Ans. 21.2% N and 24.2% S**

4. What is the percentage of trioxocarbonate (IV) ion in sodium trioxocarbonate (IV),(Na2CO3)?

**Ans.56.6% CO3 (for the CO32- ion)**

**CHAPTER 4**

**EMPIRICAL AND MOLECULAR FORMULAE (DETERMINING FORMULAE)**

These two formulas are called **determining formulas.** Knowing the percentage composition of a compound can directly lead to its empirical formula.

**An empirical formula (simplest formula) for a compound is defined as the formula of a substance that shows the simplest ratio of atoms in the compound.** The empirical formula of a substance is written with the smallest integer (whole number) subscripts. Example, the empirical formula of hydrogen peroxide (H2O2) is HO, the empirical formula of benzene (C6H6) is C2H2 (ethyne used in welding).

**Molecular formula is defined as the formula that describes the precise number of atoms of different elements in a molecule of the substance.**

*We can obtain the empirical formula from the composition of the compound by converting masses of the elements to moles. If you are given in percentages, then the masses of each element in the sample equals the numerical value of the percentage.*

***Solved Example 1***:

On analysis, a sample weighing 1.587g is found to contain 0.483g of Nitrogen and 1.104g of Oxygen. What is the empirical formula of the compound?

**Solution**

**Step 1:** Convert the masses or percentages to moles by dividing the mass or percentage by the atomic mass of each element in the compound.

**Step 2:** Then divide each mole by the smallest one.

**Step 3:** Write the mole ratio of the element in the compound to give the empirical formula

Step 1:  = 0.0345 mole of nitrogen

 = 0.06900 mole of oxygen

Step 2:mole of nitrogen;moles of oxygen

Step 3: The ratio of nitrogen atoms to that of oxygen is 1: 2. Hence, the empirical formula is NO2

Alternatively;

|  |  |  |
| --- | --- | --- |
| **Element involved** | N | O |
| **%(Mass)** | 0.483 | 1.104 |
| **Mole ratio(number of mole)** |  |  |
| **Divide through by the smallest mole ratio** |  |  |
| **Whole number** | 1 | 2 |
| **Empirical formula:NO2** | | |

***Solved Example 2***:

An analysis of sodium dichromate gives the following mass percentages: 17.5% of Na, 39.7% of Cr and 42.8% of O. What is the empirical formula of this compound?

***Solution:***

Recall: The mass of each element in the sample equals to the numerical value of the percentage. Hence, 17.5g of Na, 39.7g of Cr and 42.8g of O.

Step1: Convert the masses to moles by dividing the mass by the atomic mass of each element in the compound:

Step2: Then divide each mole by the smallest one;

Step 3: Write the mole ratio of element in the compound. The subscripts are not all integers; Na1.0Cr1.0O3.5

However, they can be made into integers by multiplying each one by 2: Na2.0Cr2.0O7. The empirical formula becomes Na2Cr2O7.

Alternatively;

|  |  |  |  |
| --- | --- | --- | --- |
| **Element involved** | Na | Cr | O |
| **%(Mass)** | 17.5 | 39.7 | 42.8 |
| **Mole ratio(number of mole)** |  |  |  |
| **Divide through by the smallest mole ratio** |  |  |  |
| **Whole number** | 1.0× 2 = 2 | 1.0 × 2= 2 | 3.5 × 2 = 7 |
| **Empirical formula: Na2Cr2O7** | | | |

***Solved Example 3***:

A sample of compound weighing 83.5g contains 33.4g of sulphur. The rest is oxygen. What is the empirical formula of the compound?

***Solution***

Mass of oxygen = mass of the compound – mass of sulphur

Mass of oxygen = 83.5g – 33.4g = 50.1g

Step 1: Convert the masses or percentages to moles by dividing the mass by the atomic mass of each element in the compound.

Step 2: Then divide each mole by the smallest one.

Step 3: Write the mole ratio of the element in the compound to give the empirical formula

Step 3: The ratio of sulphur atoms to that of oxygen is 1: 3. Hence, the empirical formula is SO3.

**Solved example 4:** A compound Y containing C , H and O only was burnt in a stream of pure oxygen. The carbon (IV) oxide and water produced were collected in pre-weighed absorption tubes and carefully re-weighed. Readings obtain are:

Initial mass of Y = 20.63mg

Mass of CO2 produced = 57.94mg

Mass of H2O produced = 11.85mg

What is the empirical formula for the compound? [H = 1, C = 12, O = 16]

**Solution**

The mass of C in the sample:

44 g of CO2 = 12 g of C

57.94 g of CO2 = ?

Cross multiplying we get;

57.94 × 12 / 44 = 15.80 mg

The mass of H in the sample:

18 g of H2O = 2 g of H

11.85 g of H2O = ?

Cross multiplying we get;

11.85 × 2 /18 = 1.32 mg

Mass of O in the sample:

20.63 - ( 15.80 + 1.32 ) = 3.51 mg

| Element | C | H | O |
| --- | --- | --- | --- |
| Mass | 15.80 | 1.32 | 3.51 |
| Mole ratio | 15.80/12 = 1.32 | 1.32/1 = 1.32 | 3.51/16 = 0.22 |
| Divide through by the smallest ratio | 1.32/0.22 = 6 | 1.32/0.22 0 = 6 | 0.22/0.22 = 1 |

**The empirical formula for the compound Y = C6H6O**

**Solved Example 5:** What is the empirical formula of a compound containing 40

**Solution**

| Element | C | H | O |
| --- | --- | --- | --- |
|  | 40.92 | 4.58 | 54.5 |
| Mole ratio | 40.92/12 = 3.4072 | 4.58/1 = 4.580 | 54.5/16 = 3.40625 |
| Divide through by the smallest ratio | 3.4072/3.40625 = 1.0003 | 4.58/3.40625 = 1.3446 | 3.40625/3.40625 = 1.00 |

We have C1H1.3O1 as the formula and next we need to convert 1.3 to whole

number by multiplying all the subscripts by 3.

C: 1.0003 × 3 = 3.0009

H: 1.344 × 3 = 4.0338

O: 1.000 × 3 = 3.000

Therefore, the empirical formula is C3H4O3

**MOLECULAR FORMULA**

Once you determine the empirical formula for a compound and you were given molecular mass, you can calculate its empirical formula weight and number of empirical formula units (n) in the molecule.

***Solved example 1***

An acetic acid (ethanoic acid) was analyzed to contain 39.9%C, 6.7% H and 53.4% O. Determine the empirical formula. The molecular mass of acetic acid was determined by experiment to be 60.0 amu. What is its molecular formula?

***Solution***

A sample of 100.0g (100%) of acetic acid contains 39.g C, 6.7% H and 53.4% O.

**Step 1:** Convert the masses or percentages to moles by dividing the mass by the atomic mass of each element in the compound.

**Step 2:** Then divide each mole by the smallest one.

**Step 3:** Write the mole ratio of the element in the compound to give the empirical formula.

.

To obtain the molecular formula,

+ × 16 = 30

The molecular formula of acetic acid is (CH2O) n

Empirical formula Empirical formula unit

Molecular formula = (CH2O)2 => C2H4O2

**Solved Example 2:**

On analysis, 1.0g sample of a hydrocarbon was found to contain 0.923gofcarbon. If the vapour density of the hydrocarbon is 39.0:

1. Determine its molecular formula
2. Name the compound

[H = 1.00, C = 12.00]

**Solution:**

The mass of hydrogen =Total mass of substancemass of carbon = 1.00.923 = 0.077

|  |  |  |
| --- | --- | --- |
| Elements involved | C | H |
| Mass | 0.923 | 0.077 |
| Mole ratio |  |  |
| Divide through by the smallest ratio |  |  |

Empirical formula = CH

Molecular formula = (Empirical formula)n

*Note: molecular mass = 2vapour density* = 239=78 g/mol

n(CH) = 78

n(121) = 78

13n = 78

n = 6

Therefore, molecular formula = 6(CH) = C6H6

ii) The compound is benzene

**Solved Example 3:** On analysis, an ammonium salt of an alkanoic acid gave 60.5% carbon and 6.5% hydrogen, If 0.309g of the salt yielded 0.0313g of nitrogen, determine the empirical formula of the salt [H = 1.00; C = 12.00; N = 14.00; O = 16.00]

**Solution:**

% N =

92.9.99999992.9

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Element | C | H | N | O |
| % | 60.5 | 6.5 | 10.1 | 22.9 |
| Mole ratio |  |  |  |  |
| Divide through by the smallest ratio |  |  |  |  |

Empirical formula = C7H9NO2

**Solved Example 4:** If the% of water in magnesium tetraoxosulphate (VI)crystals is 51.2%, what is in the formula MgSO4.nH2O

**Solution:**

Relative molecular masses: MgSO4 = 120; H2O = 18

Mole of MgSO4

Mole of H2O

Divide by the smallest mole:

Empirical formula = MgSO4.7H2O

Hence, n = 7

**Note:** When we say that the percentage of water in the crystals is 51.2%, we mean that 100g of crystals contains 51.2g of water. The remaining 48.8g is the mass of MgSO4

***Check your progress***: when 127g of copper combine with oxygen, 143g of an oxide are formed. What is the empirical formula of the oxide? **Ans** = Cu2O

**Solved example 5**

The empirical formula of a liquid B, is C2H4O. The relative molecular mass is 88. What is the molecular formula of B?

***Solution***

(C2H4O)n = Relative molecular mass

Relative molecular mass of C2H4O = 12 + 14 + 16 1 = 44

(C2H4O)n = Relative molecular mass

44n = 88

N = 2

The molecular formula of B = (C2H8O2)2 = C4H8O2

**Solved Example 6:** Analysis shows the empirical formula of a compound to be CH2O. Its relative molecular mass is 60. What is the molecular formula?

**Solution:**

Relative molecular mass = 60

Relative empirical formula of CH2O => (12 + 2 +16) = 30

(CH2O)n = relative molecular mass

(30)n = 60

n = 2

Molecular formula = (CH2O) n => C2H4O2

**Solved Example 7:** What is the molecular formula of the compound, A, which has an empirical formula C2H6O and a relative molecular mass of 46?

**Solution:** relative molecular mass of A = 46

Relative empirical formula of A =

Molecular formula => (C2H6O) n = relative molecular mass

(46) n = 46

n =1

Hence molecular formula = (C2H6O) => C2H6O.

Since the empirical formula unit is unity (1), the molecular formula is the same as the empirical formula C2H6O.

**Solved Example 8:** Determine the empirical formula of a compound made up of 28.7% K, 1.5% H, 22.8% P and 47.0% O.(H = 1,O = 1,P = 30.97,K = 39.10)

**Solution**

Step 1: Write down the elements in the compound and their percentages or grams. In a 100% sample (amu), K = 28.7g, H = 1.5g, P = 22.8g and O = 47.0g

Step 2: Find the mole by dividing the given percentage or gram by the atomic mass. In that 100g sample (amu), there are:

28.7/39.10 = 0.734 mole of K

1.5/1 = 1.5 moles of H

22.8/30.97 = 0.736 mole of P

47.0/16.0 = 2.938 moles of O

Step 3: Dividing the smallest number of moles given in step 2 to find the simplest whole number ratio:

0.734/0.734 = 1.0 of K

1.5/0.734 = 2.0 of H

0.736/0.734 = 1.0 of P

2.938/0.734 = 4.0 of O

KH2PO4 is the empirical formula.

**EXAMPLES INVOLVING GIVEN MOLECULAR MASS OR VAPOUR DENSITY**

***Suppose a hydrocarbon molecule has an empirical formula of C2H5 and a molecular mass of 58 (C = 12 ,H = 1). Deduce its molecular formula. The empirical formula mass = ( 2 x 12 ) + 5 = 29.***

***Dividing 58 by 29 gives 2. So the molecular formula is 2 x the empirical formula = C4H10***

1. ***Suppose a molecule has an empirical formula of simply CH, but a molecular mass of 78 ( C = 12, H = 1 ) .The empirical formula mass is 12 + 1 = 13. Therefore, 78/13 = 6 ,so the molecular formula is 6 x CH = C6H6. Where the empirical formula and molecular formula are different, you need extra information to deduce the molecular formula from the empirical formula.***

**Example 1:** It is found that 207g of lead combined with oxygen to form 239 g of a lead oxide. From the data, determine the formula of the lead oxide.

(Relative atomic masses: Pb = 207 and O = 16)

**Solution**

In this case, you first have to work out the amount of oxygen combined with the lead.

By simple logic from the law of conservation of mass, this is 239 – 207 = 32g

Pb;

For O; 

In atomic ratio terms, the 207 is equivalent to 1 atom of lead and the 32 is equivalent to 2 atoms of oxygen (1 x 207 to 2 x 16), so the formula is simply PbO2.

**Note: The mass of oxygen combined with the lead is deduced by subtracting the original mass of lead from final total mass of lead oxide. Lead (207) oxygen O (16).**

**Reacting mass 207g, 239 -207 = 32g not the real atom ratio. Atom ratio from mass/atomic mass values 207/207 = 1 of Pb and 32/16 = 2 of O. Work out the simplest whole number ratio simplest whole number atom ratio by oxide as PbO2 . It’s actually called lead (IV) oxide**

**Example 2:** It is found that 54g of aluminium forms 150g of aluminium Sulphide. Work out the formula of aluminium sulphide. (Relative atomic masses: Al = 27 and S = 32).

**Solution**

Amount of sulphur combined with the aluminium = 150 – 54 = 96g

For Al; 

For S; 

By atomic ratio, the 54 of aluminium is equivalent to 2 atoms of aluminium and the 96 of sulphur is equivalent to 3 atoms of sulphur.

Therefore, the atomic ratio is 2 to 3, so the formula of aluminium sulphide is Al2S3.

**Example 3:** It was found that 0.39g of carbon was combined with 4.61g of chlorine. Deduce the molecular formula of the compound. Atomic masses: C = 12 and Cl = 35.5

**Solution**

Step 1: Convert the masses or percentages to moles by dividing the mass by the atomic mass of each element in the compound.

Mole of Carbon = = 0.0325 mol

Mole of chlorine =  = 0.130 mol

Step 2: Then divide each mole by the smallest one.

For carbon =  = 1

For chlorine =  = 4

Step 3: Write the mole ratio of the element in the compound to give the empirical formula.

Therefore, the simplest ratio of 1:4 gives the empirical formula is CCl4. Its actually called tetrachloromethane.

**Example 4:** The ratio of carbon atoms to hydrogen atoms in a hydrocarbon is 1 : 2. If it's molecular mass is 56, what is its molecular formula?

**Solution**

✓ Total ratio = 1 + 2 = 3

✓ Multiply the molecular mass by the ratio of the element and divide by the total ratio:

For carbon;  = 18.666

For Hydrogen;  = 37.333

✓ Divide through by the smallest ratio:

For carbon;  = 1.0

For Hydrogen;  = 2.0

Hence, the molecular formula is CH2

**TAKE AWAY**

* + - 1. On analysis, the salt sodium sulfate was found to contain 32.4% sodium, 22.5% sulfur and 45.1% oxygen. Calculate the empirical formula of sodium sulfate. Atomic masses: Na = 23, S = 32 and O = 16. **Ans. The simplest ratio gives the empirical formula for sodium sulfate = 2 : 1 : 4, formula is Na2SO4**
      2. Analysis of hydrocarbon showed it consisted of 83.3% carbon and 16.7% hydrogen. Calculate the empirical formula of the hydrocarbon. Atomic masses: C = 12 and H =1 **Ans. Therefore the simplest ratio gives the empirical formula for the hydrocarbon = 5 : 12, so formula is C5H12. This is the simplest hydrocarbon molecule called pentane.**

3. What is the empirical formula of a hydrocarbon containing 0.160 moles of carbon and 0.640 moles of hydrogen? **Ans. CH4**

4. Vitamin C, which may or may not be effective in preventing colds, is an organic compound containing carbon, hydrogen and oxygen. Analysis of a carefully purified sample gives C = 40.9 %,H = 4.55%, O = 54.5%. From another experiment, the molar mass of vitamin C was found to be 180. From these data, determine:

a) the empirical formula of vitamin C. **Ans. C3H4O3**

b) the molecular formula of vitamin C**. Ans. C6H8O6**

5. If the empirical formula of a compound is CH2F. Determine the true formula of the compound. Ans. C2H4F2

**CHAPETR 5**

**LAWS OF CHEMICAL COMBINATION**

**Law of conservation of mass**

This states that masses of substances that you have before and after a chemical reaction remain exactly the same. This is always done in a closed system.

**Solved Example 1:** Suppose you left an iron pot on your kitchen countertop, and over time you find that the pot forms rust. If the metal pot originally had a mass of 101g and the rusted pot had a mass of 143g, what is the mass of oxygen that reacted with the iron to form the rust?

**Solution:**

Before rusting, m =101g

After rusting =143g

Before rusting After

101g 143g + Oxygen

MIron + MOxy = Mrust

MOxy = Mrust MIron

Mrus t= mass of rusted iron

MOxy = mass of oxygen

MIron = mass of iron

MOxy = 143g – 101g

MOxy = 42g

*One consequence of the law of conservation of mass is that, In a balanced chemical symbol equation, the total of relative formula masses of the reactants is equal to the total relative formula masses of the products.*

**Law of definite proportions (constant composition)**

*The law of definite proportion by mass states that the same chemical compound always contains the elements in the same proportions by mass irrespective of the source*. This law allows scientists to differentiate between a compound and a mixture. For instance, calcium trioxocarbonate (IV), CaCO3, regardless of where you get the CaO3from, the mass of Ca will be 40g,C = 12g, and O = 16 × 3 = 48g. The mass % of each of the element will always be the same.

***Solved example 1:***

The table below shows the collected sample of malachite, a mineral. Prove that the data below confirms the law of definite proportion.

|  |  |  |  |  |
| --- | --- | --- | --- | --- |
| Sample origin | Mass (g) | Cu (g) | C(g) | O (g) |
| Africa | 23.91 | 12.30 | 2.32 | 9.29 |
| Asia | 1.206 | 0.620 | 0.117 | 0.469 |
| Europe | 7.909 | 4.069 | 0.767 | 3.073 |
| America | 11.44 | 5.89 | 1.11 | 4.44 |

***Solution***

100

***Solved example 2***

2.50g of one sample A, of copper (II) oxide when completely reduced gave 2.02g of copper. 2.30g of another sample B of the same oxide prepared by a difference method, when reduced, gave 1.86g of copper. Using these data, illustrate the law of definite proportion or constant composition.

***Solution***

|  |  |  |
| --- | --- | --- |
| Sample | A | B |
| Mass of copper | 2.02g | 1.84g |
| Mass of copper (II) oxide | 2.50g | 2.30g |
| % of Cu in the oxide |  |  |
|  | = 80.80% | 80.86% |

From the above results, it is shown that the % composition of copper in each oxide is the same regardless of the method of preparation. Hence, the law of constant composition is applied.

**LAW OF MULTIPLE PROPORTIONS**

***This law states that If two elements combine to form more than one compound, the different masses of one which combines with a fixed mass of the other are in a simple whole number ratio***.

**What does the law mean?**

Let’s consider CO (Carbon Monoxide) and CO2 (Carbon Dioxide). Both of these compounds contain the same two elements (Carbon and Oxygen) but the way they combine is different. Now the atomic mass of C = 12g and O = 16g.

|  |  |
| --- | --- |
| CO | CO2 |
| 12g of C = 16g of O | 12g of C = 32g of O |
| Divide through by 12g to have 1g of the first element. | |
|  |  |

The ratio of the masses of the second element (oxygen) can always be reduced to a simple whole number. The ratio here is 1:2 of CO and CO2. I.e. .

Hence, the oxygen atoms in CO2 is twice that of CO.

**Note:** When two elements form a series of compound, the ratio of the masses of the second element that combine with 1g of the element can always be reduced to small whole numbers.

**Solved Example 3:** The masses of oxygen that combine with 1g of nitrogen to form three different compounds are 1.142g, 2.284g and 2.855g respectively. Show how these data illustrate law of multiple proportions.

**Solutions:**

|  |  |  |  |
| --- | --- | --- | --- |
| Compound | A | B | C |
| Mass of oxygen | 1.142g | 2.284g | 2.855g |
| Ratio of B to A |  |  |  |
| Ratio of C to A | to get a whole number. |  |  |

****

**TAKE AWAY**

1.5 grams of compound A reacts with 10 grams of compound B to produce a new compound C. What should be the mass of C according to law of conservation of mass? **Ans. 15g**

2. Irrespective of the source, pure sample of water always yields 88.89% mass of oxygen and 11.11% mass of hydrogen. This is explained by the law of ............. **Ans. Constant Composition**

3. Among the following pairs of compounds, the one that illustrates the law of multiple proportions is

(a) NH3 and NCl3

(b) H2S and SO2

(c) CuO and Cu2O

(d) CS2 and FeSO4

**Ans. (c)**

4.1.0 g of an oxide of A contained 0.5 g of A. 4.0 g of another oxide of A contained 1.6 g of A. The data indicate the law of

(a) Reciprocal proportions

(b) Constant proportions

(c) Conservation of energy

(d) Multiple proportions

**Ans. (d)**

5. In compound A, 1.00 g nitrogen unites with 0.57 g oxygen. In compound B, 2.00 g nitrogen combines with 2.24 g oxygen. In compound C, 3.00 g nitrogen combines with 5.11 g oxygen. These results obey the following law

(a) Law of constant proportion

(b) Law of multiple proportion

(c) Law of reciprocal proportion

(d) Dalton’s law of partial pressure

**Ans. (b)**

6. A mixture of 12 g of carbon and 64 g of sulphur is heated until reaction is complete and none of the reactants is left over. What mass of the compound, carbon (II) sulphide is formed? **Ans. 76g**

**CHAPTER SIX**

**STOICHIOMETRY**

Stoichiometry is pronounced as ***stoy – key – om – e - tree***. It is the calculation of the quantities of reactants and products involved in a chemical reaction. Stoichiometry is often based on the relationship between mass and moles, and volumes or moles and number of atoms (6.023×1023 ions/particles/atoms). It is also based on balanced chemical reactions. Stoichiometry is the branch of chemistry which tells us the quantitative relationship between reactants and products in a balanced chemical equation.

✓**Stoichiometry is simply the mathematics behind chemistry. Given enough information, one can use stoichiometry to calculate masses, vplume, moles, and percents within a chemical equation.**

**ASSUMPTIONS:**

All the reactants are completely converted to the product and there is no side reaction occurring.

When discussing about chemical reaction, the following two laws are kept in mind;

* The law of conservation of mass
* The law of definite proportions

Stoichiometry is very essential as it talks about many relationships;

* **Mass-Mass relationship**
* **Mass-Mole relationship**
* **Mass-Volume relationship**

Suppose you are given the chemical equation below to interpret:



2 moles of hydrogen + 1 mole of oxygen → s moles of water vapour.

2(2g/mol) = 4g of H2 + 1(32g/mol) of O2→ 2(18g/mol) = 36g water of water vapour

2(22.4 L) = 44.8 L of H2 + 1(22.4 L) of O2→ 2(22.4 L) = 44.8 L of water vapour

The above equation is interpreted as:

In terms of moles: 2 moles of hydrogen gas reacts (combine) with 1 mole of oxygen gas to produce (yield) 2 moles of water vapour.

In terms of mass: Here, it is very important to know the relative atomic mass of common elements or molar mass of common compounds. For the equation above, the atomic mass of H= 1, O = 16.

4g of hydrogen combines with 32g of oxygen gas to produce 36g of water vapour (1mole of water = 18g, hence 2 moles of water = 36g of H2O).

In terms of volume(L or dm3) :1mole of any substance = 22.4 dm3(standard).44.8L of H2(2 × 22.4L) combines with22.4L of oxygen(1 × 22.4L) to produce 44.8L of water (2 × 22.4L).

N/B: The above reaction obeys the law of conservation of masses, where the total mass of the reactants [2H2 (4g) + O2 (36g)] equals the total mass of the product (2H2O) 36.0g

**Using molar volume in Stoichiometry calculation involving gases**

The volume or amount of a gas at s.t.p in a chemical reaction can be calculated from:

* s.t.p condition (at standard temperature and pressure)
* Mole ratio from the balanced equation

Gases can be either product or reactants in chemical reaction, so they can be involved in Stoichiometry problems. Note very importantly that at s.t.p (0oC = 273K and 1 atoms), 1 mole of gas occupies 22.4 L.

***Solved example 1:***

What volume of oxygen is needed to completely react with 15.0g of aluminium at s.t.p?

***Solution***

4Al(s) + 3O2 (g) ==> 2Al2O3(s)

This problem relates volume with mass.

Convert 15.0 g of aluminium to volume of aluminium:

Recall that molar mass of a substance = molar volume (22.4 dm3).

Hence, 27g of Al = 22.4 dm3

From the balanced chemical equation, there are 4 moles of aluminium. That is;

4 × 22.4 dm3 = 4 × 27 g of Al

? = 15.0 g of Al

Cross multiplying we get;

4 × 22.4 dm3 of aluminium = 3 × 22.4 dm3 of Oxygen

12.44 dm3 of aluminium ?

Cross multiplying we get;

Or

Convert the given mass of Al to mole of Al:

g of Al → Mol of Al → Mol of O2 → L or dm3 I.e convert the g(mass) of Al to moles of Al using molar mass. Then convert moles of aluminium to moles of oxygen using mole ratio in the balanced chemical equation. Finally, you convert moles of oxygen to volume in litre using molar volume.

**Note:** The mole ratio of unknown is always the numerator.

**SHORTCUT:**

To find the volume of a substance given the mass in a chemical reaction simply use;

***Mass of the given substance × 1 mole/atomic mass of given substance × Mole of unknown in the balanced equation/Mole of known in a chemical equation × Molar volume/1 mole.***

**Solved example 2:** How many grams of water are formed when 1.24L of H2 gas at s.t.p completely reacts with O2?

**Solution:** Here, the problem relates mass to volume

Write a balanced chemical equation: 

Recall: At s.t.p 1 mole of a gas (any gas) = 22.4L or dm3

Convert the volume (L or dm3) of H2 to mole of H2 using 22.4L=1mole. Then convert moles of H2 to moles of H2O using mole ratios in the balanced equation. Finally, convert the mole of H2O to g of water.

**SHORTCUT:** To find the mass of a substance given the volume in a chemical reaction, simply use:

***Volume of the given substance × 1 mole/molar volume × Mole of unknown in the balanced equation/Mole of known in a chemical equation × Molar mass of unknown/1 mole.***

**CHECK YOUR PROGRESS:** What mass of Fe will react with 5.50L of O2 at s.t.p?

**SOLUTION:**

L → mole of O2 → mole of Fe → g of Fe

***It is very important that you marry the formula on how to find volume when the mass is given and vice versa.***

**MORE EXAMPLES ON STOICHIOMETRY RELATING MASS TO VOLUME**

* How many grams of pentane (C5H12) is necessary to produce 5.0L of CO2?

**SOLUTION:**

**STEPS:**

1. Write a balanced chemical equation for the reaction:
2. Convert volume of carbon dioxide to mole: 22.4L = 1 mole

5.0L = ?

1. Mole ratio: multiply the mole (0.22 mole of CO2) obtained in step(2) by mole of ratio of C5H12
2. Convert mole of pentane to grams of pentane using:

Molar mass of C5H12 = (12 × 5) + (1 × 12) = 72.0 gmol-1

Or

Using the shortcut formula;

***Volume of the given substance × 1 mole/molar volume × Mole of unknown in the balanced equation/Mole of known in a chemical equation × Molar mass of unknown/1 mole.***

Molar mass of C5H12 = (12 × 5) + (1 × 12) = 72.0 gmol-1

Reacting

* How many grams of H2O will be produced by 58.2L of CH4  at s.t.p



**SOLUTION:**

1. Balanced already
2. Convert volume of CH4 to mole: Liters of CH4 → moles of CH4

1 mole = 22.4 L

? = 58.2 L

Cross multiply;

1. Use the balanced equation to find the mole ratio: 1 mole of CH4 gives 2 moles of H2O. Therefore, 2.598 moles of CH4will give how many moles of water?

1 mole CH4 = 2 moles of water

2.598 moles of CH4 = ?

Cross multiply;

Multiply the mole (2.60 mole of CH4) obtained above by mole ratio of water to CH4 i.e. 2.598 × 2 = 5.196 moles of water.

1. Convert mole of water to grams of water using the formula below:

Molar mass of water = 1 × 2 + 16 × 1 = 18 gmol-1

Or using the formula below;

***Volume of the given substance × 1 mole/molar volume × Mole of unknown in the balanced equation/Mole of known in a chemical equation × Molar mass of unknown/1 mole.***

**CHECK YOUR PROGRESS:**

✓How many grams of hydrogen is needed to react with 44.8dm3 nitrogen of according to the reaction:



**SOLUTION:**

1. The equation is already balanced.
2. Convert volume of N2 to mole using 
3. Mole ratio: multiply the mole (2.0 mole of N2) obtained above by mole ratio of hydrogen to nitrogen. i.e.

4. Convert mole of Hydrogen (6.0 moles) to grams of hydrogen using mass = Amount (mol) × molar mass (In this case atomic mass)

Atomic mass of H2 = 1 × 2 = 2.0g

Mass of hydrogen = 6.0 × 2.0 = 12.0g of hydrogen.

Or

Simply use the formula below:

***Volume of the given substance × 1 mole/molar volume × Mole of unknown in the balanced equation/Mole of known in a chemical equation × Molar mass of unknown/1 mole.***

* 0.220g of anhydrous disodium trioxocarbonate (IV) reacted with 20 cm3 of 5.0 moldm-3 HCl and the volume of carbon dioxide collected at 23˚C and 750 mmHg pressures was 49 cm3. What is the Stoichiometry of this reaction?

**Solution**

First, we calculate the amount of CO2 :

Amount (moles) of Na2CO3 used = mass/Molar mass. The molar mass of Na2CO3 = 2 × 23 + 12 + 3 × 16 = 106 moldm-3

Amount = 0.220/106 = 0.002075 ~ 0.0021 mol

Secondly, we find the volume of CO2 collected at s.t.p using general gas equation:

P1V1/T1 = P2V2/T2

V2 = P1V1T2/P2V2T1

V2 = 750 × 49 × 273/760 × 296 = 44.6 cm3 or 0.0446 dm3

Finally, we find the mole of CO2:

At s.t.p, 22.4dm3 of CO2 = 1 mole

0.0446 dm3 of CO2 = ?

Cross multiplying we get;

1 mole multiplying 3/ 22.4 dm3 = 0.001991 ~ 0.0020 mole

Hence, 0.0021 mole of Na2CO3 produced 0.0020 mole of CO2

i.e 1 mole of Na2CO3 produced 1 mole of CO2

The balanced equation becomes;

Na2CO3 + 2HCl => 2NaCl + CO2

**AMOUNTS OF SUBSTANCES IN A CHEMICAL REACTION (mole):**

A mole (mol) is defined as the quantity (amount) of a given substance that contains as many molecules as the number of atoms in exactly 12g of carbon-12. The number of atoms in a 12g sample of carbon-12 is called Avogadro’s number (NA) which is given as 6.023 × 1023 atoms, molecules, ions or particles.

* Mole is the amount of substance which contains 6.023 × 1023 atoms, molecules, ions or particles.

Mole = or

**THE MAGIC FORMULA:**

* **1 mole of any substance = molar mass (compound) or atomic mass (element) = Avogadro’s number (6.023 × 1023 atoms, molecules, ions or particles) = 22.4dm3 = Q/F**

The above magic formula shows that you can actually calculate any of the parameter if one of them is given. (You can relate one parameter with another).

Where n = number of moles, M.M = molecular mass, C = concentration (moldm-3), V = volume (cm3), N = number of molecules/particles/ions/atoms and = Avogadro's number.

* Mass =

**SOLVED EXAMPLES:**

1. What is:

(i)The mass in grams of a chlorine atom (Cl)?

(ii)The mass in grams of a hydrogen chloride molecule, HCl?

**SOLUTION:**

(i)

(ii) The molecular mass of HCl = 1.0 + 35.5 = 36.5g

2. How many moles of ethanol are in 10.0g of ethanol?

**SOLUTION:**

Mole =

Molar mass of C2H5OH = 12 × 2 + 1 × 6 + 16 × 1 = 46.0 gmol-1

Mole =  = 0.217 mol of C2H5OH

Alternatively,

46.0g C2H5OH = 1 mol of C2H5OH

10.0g of C2H5OH = ?

Cross multiply;

**3:** What is the mass of 2 moles of aluminium? [Al=27]

**Solution**:

Mole =

Mass = mole × Atomic mass

Mass = 2 × 27 = 54g of Al.

**4:** Zinc Iodide, ZnI2 can be prepared by the direct combination of elements. A chemist determines from the amount of element that 0.0654 mole can be formed. How many grams of Zinc Iodide is this?(ZnI2 = 319g/mol)

**Solution:**

1 mole of ZnI2 = 319g ZnI2

0.0654 mole of ZnI2 = ?

Or

Mass = mole × Molar mass

Mass = 0.0654 × 319 = 20.9 g ZnI2

1. How many moles are contained in 45.6g of lead (II) chromate?

**Solution:**

The molar mass of PbCrO4 = 107 + 52 + (16 ×4) = 323g of PbCrO4

1 mole of PbCrO4 = 323g of PbCrO4

? = 45.6g of PbCrO4

Cross-Multiplying;

Or simply use; Mole =

1. How many moles are there in 6 dm3 of oxygen gas at s.t.p.?

**Solution**

=> 0.26 moles of oxygen gas.

1. In the analytical laboratory of a pharmaceutical company a laboratory assistant was asked to make 250 cm3 of a 2.0 x 10-2 moldm-3(0.02M) solution of paracetamol (C8H9NO2). How much paracetamol should the laboratory assistant weigh out to make up the solution?

**Solution**

Relative Molar mass of paracetamol = 151

Moles = molarity x volume in dm3

Convert 250 cm3 to dm3: 250/1000 = 0.25 dm3

Mole of paracetamol required = 2.0 x 10 -2 x 0.25 = 5.0x10-3 (0.005) mol

= x = 5.0 x 10 -3 x 151 = 0.755g

1. What is the mole number of glucose (C6H12O6) in 25% (w/w) 1 kg of its aqueous solution?

**Solution**

Molar mass (M) of glucose (C6H12O6) = [(12x6) + (1x12) + (16x6)] =180gmol-1

Mass (m) of glucose = 1kg = 1000g

1. Chalcaolpyrite [CuFeS2] is an ore of copper, what is the amount of copper in 3710kg?[Cu = 63.5,Fe = 55.8, S = 32]

**Solution**

Molar mass of CuFeS2 = 63.5 + 55.8 + (322) = 183.3.

Mass of Chalcaolpyrite = 3710kg = 371000g

**Converting Number Of Molecules Or Atoms To Mass And Vice Versa**

**Solved example 1:** How many grams of zinc contain:

1. 6.0 × 1023 atoms
2. 6.0 × 1020 atoms

(Zn = 65.0)

**Solution**

1. Recall: Atomic mass of an element = NA (6.02 × 1023atoms)

Hence, 65.0g of zinc contains 6.0 × 1023 atoms

1. 65.0g of zinc = 6.02 × 1023 atoms

? = 6.0 ×1020 atoms

Cross multiplying, we get;

**Solved example 2:** How many molecules are there in a 3.46g sample of hydrogen chloride?

**Solution**

The molar mass of HCl = 1+ 35.5 = 36.5. (*Molar mass of HCl)*

36.5g HCl =

3.46g HCl =?

Cross multiplying, we get;

***Check your progress***

Hydrogen cyanide is a highly poisonous substance. How many molecules are there in 0.56g of HCN? **Ans. 1.24859 × 1022 molecules.**

**Converting Mass Of Reactant To Mass Of Product**

We can calculate the mass of a substance produced if the mass of reactant is given and vice versa.

**Solved example1:** What mass of magnesium oxide is produced from the complete combustion of 12g of magnesium?

**Solution**

We write a balanced chemical equation for the reaction:

The given question relates mass of magnesium with mass of magnesium oxide. Hence, from the above balanced equation;

i.e

0

**Solved example 2:** During the extraction of iron from haematite (Fe2O3) using carbon monoxide, how many grams of iron can be produced from 1000g Fe2O3?

**Solution**

We write a balanced chemical equation for the reaction:

160g 2

160g

1000g ?

Cross multiply;

Hence, 1000g produces 700g Fe.

***Check your progress***

What mass of zinc tetraoxosulphate (VI) can be obtained from the reaction of 10g of zinc with an excess of dilute tetraoxosulphate (VI) acid? **Answer = 24.8g ZnSO4**

**Solved example 3:** If 4.2g of sodium hydrogen trioxocarbonate (IV) are heated, what mass of anhydrous sodium trioxocarbonate (IV) will be formed?

**Solution**

First, we write a balanced chemical equation:

?

of NaHCO3

1. Calculate the mass of carbon (IV) oxide produced by heating 15g of limestone.

**SOLUTION:**

First, we write a balanced chemical equation:

The molar mass of CaCO3 = 40 + 12 + 16 × 3 = 100g/mol

The molar mass of CaO = 40 + 16 = 56g/mol

The molar mass of CO2 = 12 + 16 × 2 = 44g/mol

This question relates CO2 with CaCO3.

100g of CaCO344g of CO2

15g of CaCO3?

This means if 100g of CaCO3 produces 44g of CO2, how many will be produced by 15g of CaCO3?

Cross multiply to get the mass of CO2 produced:

Hence, 6.6g of CO2 will be produced.

**We can also relate quantities of two reactants or products. E.g today, Cl is produced from NaCl by electrochemical decomposition. Formerly, Cl was produced by heating HCl with MnO2.**

How many grams of HCl react with 5.0g of MnO2 according to the above equation?

**SOLUTION:**

**h h**

**CHECK YOUR PROGRESS:**

How many gram of oxygen gas will combine with 5.0 × 103g of zinc sulphide to produce zinc oxide and sulphur (IV) oxide?

**Converting Reacting Volumes to Mass:**

Calculations involving reacting volumes of gases are very simple as they depend on facts that a mole of a gas occupies 22.4dm3.

**Example 1:** What volume of carbon (IV) oxide at s.t.p is produced by burning 12g of carbon?

**Solution**

Balanced chemical equation: C(s) + O2(g) => CO2(g)

Recall: 1 mole of a substance = atomic mass or molar mass =22.4dm3.

The above question relates atomic mass of carbon with volume of CO2. Hence, 12g of carbon =22.4dm3.

Or

***Mass of the given substance × 1 mole/atomic mass of given substance × Mole of unknown in the balanced equation/Mole of known in a chemical equation × Molar volume/1 mole.***

**Example 2:** What volume of hydrogen at room temperature and pressure is evolved when 0.32g of zinc reacts with dilute hydrochloric acid?

[Volume at room temperature and pressure (r.t.p) = 24.0dm3, Zn = 65]

**Solution**

Balanced chemical equation: Zn(s) +2HCl(aq) =>H2(g) + ZnCl2(aq)

From the above balanced equation,1 mole of Zn gives 1 mole of hydrogen.

65gZn 24dm3 H2 at r.t.p

0.32g Zn ?

***This means that if 65g of zinc produces 24 dm3 of hydrogen, what volume of hydrogen will be produced by 0.32g of zinc?***

**Example 3:** If 4.5g of liquid **A** vapourise to give 1.12dm3 of vapour at s.t.p, what is the relative molecular mass of **A**?

**Solution**

1.12dm3 4.5gof **A**

22.4dm3?

***This means that if 1.12dm3of vapour are occupied by 4.5g of A, 22.4 dm3(standard volume) will occupy how many grams?***

**Example 4:** What volume of hydrogen is formed when 12g of magnesium react with an excess dilute acid?

**Solution**

First, write the equation for the reaction of magnesium with any dilute acid:

Mg(s)+ 2HCl(aq)=>MgCl2(s)+H2(g)

1 mole: 2moles1mole : 1mole

The mole of Mg in 12gofthemetal:

Amount (mol) =

Amount =

The volume of hydrogen produced from 0.5 mole of Mg:

1 mole of Mg = 22.4dm3

0.5 mole of Mg = ?

**Example5:** What is the volume of carbon dioxide obtained by heating 10 g of calcium trioxocarbonate (IV)?

**Solution**

1mole of CaCO3 1mole CO2

Mole of

of

In the balanced equation;1 mole of produces 1mole of CO2.

1 mole of = 22.4dm3

* 1. Mole of =?

Since they have the same number of moles, the volume of CO2 obtained is.

***Check your progress***

Aluminium metal reacts vigorously with aqueous tetraoxosulphate (VI) acid to produce aqueous aluminium tetraoxosulphate (VI) and hydrogen gas:

Determine the volume of hydrogen produced at s.t.p when a 2.0g piece of aluminium completely reacts. **Answer: 2.49 dm3 of hydrogen.**

**Example 6:** How many moles of oxygen are in 2.71 x 1025 molecules of carbon dioxide?

**Solution**

The formula for carbon (IV) oxide is CO2.

From the formula, there are two atoms of oxygen.

6.02 × 1023 molecules = 2 atoms of oxygen

2.71 × 1025molecules = ?

**LIMITING AND EXCESS REACTANTS**

***Limiting reactant or Limiting reagent is a reagent that is completely used up or consumed in the reaction.*** Therefore, it limits the reaction from continuing. ***The excess reagent is the reactant that could keep reacting if the other had not been consumed (reactants that are not used up when the reaction is finished are called excess reagents).***

For many real- life chemical reactions, one or more of the reactants is limited in quantity. For example, the petrol reacts with oxygen from the air and while the engine runs, forming products such as carbon monoxide and water, as well as some others. After the litre of petrol is all used up, there is still plenty of oxygen in the air. In fact, after a whole tank full of petrol is burned, there is still up plenty of oxygen in the air. The limiting reactant in this case is the petrol, since it is completely consumed in the reaction.

**Example1**: For the reaction C + O2 ==> CO2. If 3 g of carbon are available when there is plenty of oxygen.

a) Which reactant is the limiting reactant?

b) How many grams of CO2 could be produced?

**Solution**

C + O2 ==> CO2

a) Carbon is the limiting reactant because it finishes as the reaction goes to completion.

b) Mole of carbon = Reacting mass/Atomic mass

1mole of C =1 mole of CO2 (from balanced chemical equation)

0.25 mole of C=?

Cross multiplying;

Molecular mass of CO2 = 12+16 × 2 = 44g

44g of CO2 = 1mole of CO2

? = 0.25 mole of CO2

Cross multiplying;

**That is, if the volume is in dm3 (litres) at room temperature and pressure, or Vgas = 22.4 x moles of gas.**

**If the gas volume is given in cm3, then dm3 = V/1000 (since 1dm3 =1000cm3).**

For instance, given the equation: HCl(g)+NH3(g) ===>NH4Cl(s);

1 mole hydrogen chloride gas combines with 1 mole of ammonia gas to give 1 mole of ammonium chloride solid, since from Avogadro's law, equal volumes of gases at the same T&P, have the same number of molecules and equal numbers of moles have the same number of molecules, and this gives rise to Gay-Lussac's Law of combining volumes, and we can then logically say directly from the equation, 1 volume of hydrogen chloride will react with 1 volume of ammonia to form solid ammonium chloride.

So, if 50cm3 HCl reacts, you can predict 50cm3 of NH3will react.

**TAKE AWAY**

* + - 1. Only 4.01 of methane are available for the reaction:

CH4 + 2O2 ==> CO2 + H2O. There is plenty of oxygen.

a) Which reactant is the limiting reactant?

b) How many grams of H2O will be produced?

**Ans. a) CH4 b) 9.01g of H2O**

2. How many molecules of oxygen would occupy a volume of 2.24 cm3 at s.t.p? [Molar volume at s.t.p. = 22400 cm3, Avogadro's number = 6.02 × 1023]. **Ans. = 6.02 × 1019 molecules of oxygen**.

3. How many moles of copper are in 6,000,000 atoms of copper? **Ans. 9.9 × 10-19 moles of copper**

4. How many atoms are in 5 moles of silver? **Ans. 3.01 ×1024 atoms of silver**

5. How many atoms of gold are in 1 gram of gold? **Ans.3.06×1024 atoms of gold**

6. How many moles of lithium (Li) are in 1 mole of lithium hydride (LiH)? **Ans. 6.1moles of lithium**

7. How many moles of oxygen are in2.71x1025 molecules of carbon dioxide (CO2)? **Ans. 90.03 moles**

8. How many atoms of hydrogen are in 1 mole of water (H20)? **Ans.1.20 × 1023 atoms of hydrogen**

9. 24g of magnesium reacts with dilute hydrochloric acid. Calculate the volume of hydrogen gas evolved at s.t.p.

[Mg = 24, Molar volume of the gas at s.t.p = 22.4dm3]. **Ans. 22.4dm3 of hydrogen gas**

10. Given the equation: N2(g)+ 3H2(g)==>2NH3(g)

1. What volume of hydrogen reacts with 50cm3 of nitrogen?
2. What volume of ammonia will be formed?

**Ans. A) 150cm3 B) 100 cm3**

11. What volume of oxygen in dm3 would be required to burn completely 8.0g of methane according to the equation below?

CH4(g) + 2O2(g) → CO2(g) + 2H2O(g) [Molar volume = 22.4 dm3, C = 12, H = 1, O = 16]

**Ans. 22.4 dm3**

12. A 0.84 g of aluminium reacted completely with chlorine at standard temperature and pressure.

i. Write an equation for the reaction.

ii. Calculate the volume of chlorine used

[Al = 27.0, Cl = 35.5, Molar volume of gas = 22.4 dm3]

Ans. i. 2Al(s) + 3Cl(g) 2AlCl3

* + 1. 1.05 dm3

**CHAPTER 7: GAS LAWS**

**INTRODUCTION**

The gas laws are a series of mathematical relationships that relate the following four variables:

T = Temperature

P = Pressure

V = Volume

n = Moles

The above four variables describe a gas.

**‌Temperature**

I. The temperature of a gas determines the average kinetic energy of the particles.

2. While the average kinetic energy of a collection of gases at a given temperature will be same, the velocity at which they travel will not. This is because the mass values of the various gases are different.

3. Temperature is usually measured in either Celsius or Kelvin. Both are related to one another:

K = °C + 273

**‌Volume**

Gas volumes are expressed in four different ways: Litre - L, milliliter - mL, centimeter cube - cm3, and decimeter cube - dm3.

**‌Pressure**

I. The force that the gas particles exert over a unit area:

2. Pressure is a measure of the total force exerted by the moving particles of a gas as they collide with the walls of the container.

3. Atmospheric pressure is measured in terms of: Atmosphere (atm).

1 atm = 760mmHg = 760torr = 101325N/m2 = 1.01×105Pa(Pascal)

Example: Convert 0.89atm to torr.

=>1 atm = 760 torr

0.89 atm = ?

**‌Standard Temperature and Pressure (STP)**

In order to study the effects of changing temperature and pressure on a gas, one must have a standard for comparison. STP represents a pressure and temperature that are fairly easy to reproduce.

* **Factors Affecting Pressure**

**Amount of Gas:** Increasing the number of particles increases collisions, which increases pressure. Removing particles reduces pressure.

**Volume:** Increasing the volume will decrease the pressure of a gas since collisions are less likely. Decreasing the volume has the opposite effect.

**Temperature:** Increasing the temperature increases the speed of the molecules, which leads to more collisions and greater pressure. Decreasing the temperature has the opposite effect.

**Charles' Law**

**Charles’ law states that the volume of a given mass of a gas is directly proportional to its absolute temperature in Kelvin provided that the pressure remains constant**.

1. Charles’ law relates volume and temperature.

2. As the temperature decreases, the volume of a gas decreases. As temperature increases, the volume of a gas increases.

3. The volume of a gas at constant pressure is directly proportional to the absolute temperature.

4. Charles' Law helped proven the existence of absolute zero.

5. Mathematically, this is:

**Example 1:** A sample of gas occupies 24.0m3 at IOO.OK. Determine its volume at 400.0K.

**Solution**

**Example 2:** Gas in a balloon occupies 2.50 L at 300.0K. At what temperature will the balloon expand to 7.50L?

**Solution**

**Example 3:** A given mass of a gas occupied 150 cm3 at 25 and a pressure of 1.013 × 105 Nm-2. Calculate the temperature at which its volume will be doubled at the same pressure.

**Solution**

**Note:** The same pressure means at constant pressure. Hence, Charles’ law is applied.

300 × 300 = 150

**‌Boyle's Law**

**This states that the volume of a given mass of a gas is inversely proportional to the pressure provided that the temperature remains constant**.

I. The volume of a gas at constant temperature is inversely proportional to the pressure.

2. Like Charles' Law, we can write:

P1V1 = P2V2

**Example 1:** The gas in a balloon has a volume of 4.00L at IOO.O kPa. The balloon is released into the atmosphere and the gas in it expands to a volume of 8.00L. Determine the pressure on the balloon at the new volume.

**Solution**

P1V1 = P2V2

100 × 4.0 = P2 × 8.0

P2 = 50.0 kPa

**Example 2:** If the pressure of a 2.50m3 sample of gas is 1.50 atm, what volume will the gas occupy if the pressure is changed to 7.50 atm?

**Solution**

P1V1 = P2V2

1.50 × 2.50 = 7.50 × V2

**‌COMBINEDGASLAW**

1. The combined gas law states the relationship among pressure, volume and temperature of a fixed amount of gas:

The volume occupied by a given amount of gas is directly proportional to the absolute temperature divided by the pressure.

2. The combined gas law allows one to work out problems involving more variables that change.

**Example 1:** A gas at 11O.0 kPa and 30.0°C fills a flexible container with an initial volume of 2.00L. If the temperature is raised to 80.0°C and the pressure increased to 440.0 kPa, what is the new volume?

**Solution**

T1 = 273 + 30 = 303K, T2 = 273 + 80 = 353K, P1 = 110.0KPa, P2 = 440.0kPa, V1 = 2.0L, V2 = ?

0.582L

**Example 2:** At O.O°C and 1.00 atm pressure, a sample of gas occupies 30.0mL. If the temperature is increased to 30.0°C and the entire gas sample is transferred to a 20.0mL container, what will be the gas pressure inside the container?

**Solution**

T1 = 273 + 0 = 273K, T2 = 273 + 30 = 303K, P1 = 1.0 atm, P2 = ?, V1 = 30.0L, V2 = 20.0L

**Example 3:** If 200 cm3 of carbon (iv) oxide were collected over water at 18 and 700 mmHg, determine the volume of the dry gas at s.t.p. [Standard vapour pressure of water at 18 = 15 mmHg].

**Solution**

Pressure of the dry gas (P1) = 700 – 15 = 685 mmHg, V1 = 200 cm3, T1 = 18 273 = 291K, P2 = 760 mmHg, T2 = 273K

**‌Gay-Lussac's Law and Avogadro's law**

**Gay-Lussac's law states that when gases react they do so in volumes which are in simple ratios to one another and to the volumes of the products if gaseous provided temperature and pressure remain constant. Or Gay-Lussac’s Law states that, at constant volume, the pressure of a gas is proportional to its absolute temperature, specified in Kelvin.**

For example, if you double the temperature of a gas, you double its pressure, and vice versa.

Gay-Lussac’s Law can be expressed as the equation:

P1T2 = P2T1For example, if we increase the temperature of a 7.5 mL gas sample at atmospheric pressure from 293.15 K (20.00°C) to 373.15 K (100.00°C), we can calculate the pressure at the higher temperature by substituting the known values in the Gay-Lussac’s Law equation:

(101,325 Pa)(373.15 K) = (x Pa)(293.15 K)

Solving for x, we find that the gas pressure at the higher temperature is about 128,976 Pa. This law applies to only gases not solids or liquids. For instance, hydrogen gas reacts with oxygen gas at 100°C to form steam according to the equation below:

2H2(g) + O2(g) => 2H2O(g)

2mol 1mol 2mol

2vol 1vol 2vol

2 cm3  1cm3  2cm3

From the above reaction, it means that when water is in its gaseous state, 2 vol of H2 will combine with 1 vol of oxygen to produce 2 vol of H2O to give a simple mole ratio of 2:1:2

**SOLVED EXAMPLE 1**

65 cm3 of H2 are spark of with 30 cm3of O2 at 100°C and 1 atm. Calculate the total volume of residual gas at:

a) 100°C

b) 37°C (room temperature)

**Solution**

2H2(g) + O2(g) => 2H2O(g)

2 : 1 2

2vol 1vol 2vol

Before sparking: 65cm3 30cm3 ........

On sparking: 60cm3  30cm3 60cm3

After sparking: 5cm3 0cm3  60cm3

Total volume of residual gas = 5cm3 of hydrogen gas left + 60cm3 of the steam formed.

**Total volume of residual gas = 65cm3 of gases.**

**Explanation:**

1. From stoichiometry, 2 volumes of H2 will combine with 1 volume of O2 to produce 2 volumes of H2O (steam). Going by the volumes of gases available, what volume of O2 will you need to completely spark 65cm3 of H2?

1/2 of 65 = 1/2 × 65 = 65/2 = 32.5cm3 of Oxygen.

But the volume of O2 available is only 30 cm3 (you can't use the volume you don't have) implying that it is enough to completely spark the available H2. So we make use of the available O2 which means that 30 × 2 = 60cm3 will react while 65 - 60 cm3 gives 5 cm3 of H2 that will be left unreactive. Since, the O2 is not enough to completely react with H2. It means it was used up (limited reagent) and nothing was left unreactive.

b) Recall that; no steam was produced before the sparking. Therefore, 60cm3 of steam will be left as a residual gas in addition to the 5 cm3 of the unreactive H2 to give 60 + 5 = 65cm3 of residual gas.

**SOLVED EXAMPLE 2**

25cm3 of CO reacts with O2 to give CO2. How many volume of O2 will combine to produce CO2.

**Solution**

2CO + O2 => 2CO2

2mol 1mol 2mol

2vol 1vol 2vol

25cm3 vol 25cm3

= 25cm3/2 = 12.5cm3

**SOLVED EXAMPLE 3**

25cm3 of H2 was sparked with 20cm3 of O2 at 120°C. What is the total volume of the residual gas?

**Solution**

2H2(g) + O2(g) => 2H2O(g)

2: 1 : 2

Before sparking: 25cm3 20cm3 ........

On sparking: 25cm3 12.5cm3 25cm3

After sparking: 0cm3 7.5cm3 25cm3

**Total volume of the residual gas = 25 + 7.5**

**= 32.5cm3**

**SOLVED EXAMPLE 4**

35cm3 of H2 was sparked with 12cm3 of O2 at 110°C. What is the total volume of gas left after the reaction?

**Solution**

2H2(g) + O2(g)=> 2H2O(g)

2: 1: 2

Before sparking: 35cm3 12cm3 .......

On sparking: 24cm3 12cm3 24cm3

After sparking: 11cm3 0cm3 24cm3

**Total volume of the residual gas = (11 + 24) cm3 = 35cm3**

**SOLVED EXAMPLE 5**

10 cm3 of H2 was sparked with 10cm3 of O2 at s.t.p. Calculate the volume of the unreacted gas.

**Solution**

2H2(g) + O2(g) => 2H2O(g)

2 : 1 : 2

Before sparking: 10cm3 10cm3 ........

On sparking: 10cm3 5cm3 10cm3

After sparking: 0cm3  5cm3 10cm3

**Volume of unreacted gas = 5 cm3**

**Example 6:**

50 cm3 of methane was sparked with 150 cm3 of air containing 21%of oxygen. Calculate the volume of the residual gases if methane combines with oxygen as follows:CH4(g) + 2O2(g) →2H2O(g) + CO2(g)

**Solution**

We begin by calculating the volume of oxygen by obtaining the value of 21% of 150 cm3 of air, i.e. V = 21/100 × 150 = 31.5 cm3 CH4(g) + 2O2(g) → 2H2O(g) + CO2(g)

Volume ratio: 1 : 2 : 2 : 1Volume available: 50 cm3 31.5 cm3 0cm3 0cm3Reacting volume: 15.75 cm3 31.5cm3 31.5 cm3 15.75 cm3Residual volume: 34.25 cm3 0 cm3 31.5cm3 15.75 cm3

Finally, we now calculate the total volume of residual gases as follows. V = remaining volume of air + excess reactants + volume of products = 118.5 cm3 + 34.25 cm3 + 0cm3 + 31.5cm3 + 15.75cm3 = 200 cm3

**SOLVED EXAMPLE 7**

What volume of hydrogen will be leftover when 300cm3 of oxygen and 1000cm3 of hydrogen are exploded?

2H2(g) + O2(g) => 2H2O(g)

2: 1: 2

Before sparking: 1000cm3 300cm3 ………

On sparking: 600cm3 300cm3 600cm3

After sparking: 400cm3 0cm3 600cm3

**Volume of hydrogen gas leftover = 400cm3**

**AVOGADRO'S LAW**

**Avogadro's Law states that 'equal volumes of gases at the same temperature and pressure contain the same number of molecules' or moles of gas.**

I. The pressure of a given gas varies with the Kelvin temperature when the volume remains constant.

2. This is expressed as:

As the temperature of an enclosed gas increases, the pressure increases at constant volume.

**Example 1:** The gas in a used aerosol can is at a pressure of 103kPa at 25°C. If the can is thrown onto a fire, what will the pressure be when the temperature reaches 928°C?

**Solution**

To get from °C to K, add 273.

**Example 2:** The pressure of a gas in a tank is 324.24kPa at 295.0K. Determine the gas pressure if the temperature is raised to 333.0K.

**Solution**

**Example 3:** A gas in a sealed container has a pressure of 125.0kPa at 30.0°C. Determine the temperature in the container if the pressure is increased to 201.0kPa.

**Solution**

**DALTON'S LAW OF PARTIAL PRESSURE**

**This law states that the total pressure in a gas mixture is the sum of the partial pressures of the individual components.**

* The partial pressure of a gas is the pressure of an individual gas in a gas mixture that contributes to the total pressure of the mixture.

This is expressed mathematically as:

PTotal = PT = P1 + P2 + P3 +.....

Where Pgas =partial pressure for the gas in question, ngas =moles of the gas in question, nTotal = total moles of all gases in the mixture and PTotal = total pressure.

***Note:*** *The formula for partial pressure involves multiplying the mole fraction by the total pressure. The mole fraction is simply the number of moles of a particular gas divided by the total moles of gas in the mixture.*

Mole fraction can also be determined as:

a)

b)

* Dalton's law is often used to determine the pressure of a gas collected overwater.

**Example 1:** Hydrogen gas is collected overwater at a total pressure of 95.0kPa. The volume of gas collected is 28.0mL at 25.0°C. Determine the partial pressure of the hydrogen gas if the water vapor pressure is 3.17kPa.

**Solution**

PTotal = PH2 + PH2O

95.0 = PH2 + 3.17

PH2 = 95.0 - 3.17

PH2 = 91.8kPa

**Example 2:** A gas is collected over water at 50.0°C and a barometric pressure of 105.00kPa. Determine the pressure of the gas if the water vapor pressure is 12.34kPa.

**Solution**

PTotal = Pgas + PH2O

105.0 = Pgas + 12.34

Pgas = 105.0 - 12.34

Pgas = 92.66kPa ~ 92.7kPa

**Example 3:** Air contains O2, N2, CO2, and trace amounts of other gases. What is the partial pressure of O2 at 101.30kPa of total pressure if the partial pressures of N2, CO2, and other gases are 79.10kPa, 0.040kPa, and 0.94kPa, respectively?

**Solution**

PTotal = PO2 + PN2 + PCO2 + POTHERS

101.30 = PO2 + 79.10 + 0.040 + 0.94

101.30 = PO2 + 80.08

PO2 = 101.30 - 80.08

PO2 = 21.22kPa

**Example 4:** A container has three gases mixed together at STP. The container has (by volume) 10% He, 40% neon, and 50% argon. What is the partial pressure of the three gases?

**Solution:**

Because this mixture of gases is at STP, the total pressure will be standard pressure, 760torr. Because the percentages of the gases have been given, you can setup the following:

He 10% (0.10)(760) = 76 torr

Ne 40% (0.40)(760) = 304 torr

Ar 50% (0.50)(760) = 380 torr

Check your work. Does the sum of the partial pressures equal the total pressure of 760 torr?

**Example 6:** Calculate the mole fraction of benzene in solution containing 30% by mass in CCl4.

**Solution**

Concentration 30% by mass shows that 30g C6H6 is dissolved in 100g of solution or 70g CCl4. So calculate the moles of each by dividing their masses with their molecular masses. Then calculate the moles fraction by using the formula:

For 100g of the solution:

Mass of benzene = 30g

Mass of carbon tetrachloride = 100 - 30 = 70g

Molar mass of C6H6 = 12 × 6 + 6 × 1 = 78gmol-1

Molar mass of CCl4 = 12 + (35.5 × 4) = 154gmol-1

**Example 6:** A mixture of gases exerted a pressure of 840 torr consists of 0.2 mole of oxygen, 0.3 mole of hydrogen and 0.7 mole of carbon (IV) oxide. Determine the partial pressure contributed by:

a) Hydrogen

b) Carbon (IV) oxide

Calculate the mole fractions of the gases in the mixture above.

**Solution**

nTotal = 0.2 + 0.3 + 0.7 = 1.2moles in the mixture.

Pressure contributed by hydrogen in the mixture is calculated as follows:

Pressure contributed by carbon (IV) oxide in the mixture is calculated as follows:

The mole fractions of the gases are calculated as:

Hence, the mole fractions of all gases in the mixture add up to 1.0 (0.250 + 0.583 + 0.167).

**The sum of mole fraction of all components of a solution is unity.**

**‌IDEAL GAS LAW**

The number of moles is the fourth variable that can be used along with pressure, volume and temperature to describe a gas.

The ideal gas law is: PV = nRT

Where P = Pressure, V= Volume, n = Number of moles, R = Molar constant and T = Temperature.

If pressure is expressed in atmospheres, then:R = 0.0821L.atm/mol.K

If pressure is expressed in kPa, then:R = 8.314L.kPa/mol.K

**Example 1:** Determine the number of moles of gas in a 3.00L container at 300.0K and a pressure of 1.50atm.

**Solution**

PV = nRT

**Example 2:** Determine the Celsius temperature of 2.49 moles of gas contained in a 1.OOL container at a pressure of 143kPa.

**Solution**

PV = nRT

K = °C + 273

°C = K - 273

°C = 6.91 - 273

°C = - 266°C

**Example 3:** Determine the molecular weight of a gas if 4.50g of it occupies 4.0L at 950 mmHg and 182. [R = 0.082057L.atm.mol-1K-1]

**Solution**

Mass (m) = 4.50g, P = 950 mmHg, R = 0.082L.atm.mol-1K-1,V = 4.0L, T = 182 = 455K, n = ?

Convert the unit of pressure from mmHg to atmosphere (atm) to be in line with the unit of R.

760 mmHg = 1 atm

950 mmHg = ?

Using the ideal gas equation, PV = nRT, we can calculate the value of n. After that, the molecular weight can then be determined using:

From ideal gas equation,

**GRAHAM'S LAW OF DIFFUSION**

**Diffusion:** This is the tendency of molecules to move toward areas of lower concentration until the concentration is uniform throughout.

**Effusion:** This is the tendency of a gas to escape through a tiny hole in a container.

* **Gases of lower molar mass diffuse and effuse faster than gasses of higher molar mass.**

Graham’s law of diffusion states that the rate of diffusion of a gas at a given temperature is inversely proportional to the square root of its molecular mass, m.

For two gases with molecular masses m1 and m2; and rates of diffusion R1 and R2:

Comparing the rates of diffusion of equal volumes of two gases in times t1 and t2, we have:

* The lighter or less dense a gas is, the greater is its rate of diffusion.
* If the gases are at the same temperature and pressure, they must have the same molar volumes. Therefore,

By the above last equation, equation (5), you could determine the value of any of the parameters by establishing an equation with those whose values are given.

Note: The vapour density (v.d.) of a gas is equal to half its relative molecular mass (r.m.m.).

That is molecular mass = 2 x v.d

**Example 1:** 30cm3 of a gas, the empirical formula of which was CH3, diffuses through a porous partition in 45.2 seconds. 30cm3of hydrogen diffused in 11.7 seconds under the same conditions. Calculate:

1. the vapour density of the gas
2. the molecular formula of the gas

**Solution:**

(i).The two gases are of equal volumes,

.

(ii) xCH3 = 30

(12 + 1 × 3)x = 30

15x = 30

X = 2

**Example 2:** If 80cm3 of methane (CH4) diffuses through a porous membrane in 20 seconds while 50cm3 of gas Q diffuse through the same membrane in 25 seconds, calculate the molecular mass of gas Q.[CH4 = 16.0]

**Solution**

Square both sides

**Example 3**

Calculate the molar mass of a given gas whose diffusion rate is 2.92 times the diffusion rate of ammonia.

**Solution**

We know that the diffusion rate is 2.92 times that of ammonia, therefore we understand that the ratio of diffusion rate of the given gases should be

The molar mass of NH3 = 14 + 1 × 3 = 17g/mol.

Square both sides;

M1 = 1.993 ~ 2

M1 = 2.0

**Example 4:** If equal amounts of helium and argon are placed in a porous container and allowed to escape, which gas will escape faster and how much faster?

**Solution:**

1. Set rates and get atomic weights:

rate1 = He = x

rate2 = Ar = 1

The atomic weight of He = 4.00

The atomic weight of Ar = 39.95

1. Graham's Law is:

Substituting, we have:

Helium escapes faster than Ar. It does so at 3.16 times the rate of the argon.

**Example 5:** What is the molecular weight of a gas whose ion rate is 1/50 as fast as hydrogen?

**Solution:**

1. Set rates and molecular weights:

Rate1 = unknown gas = 1

Rate2 = H2 = 50

Remember, 1/50th of 50 is1.

The molecular weight of H2 = 2.0

The molecular weight of the other gas = x.

1. By Graham's Law, we have:

1Square both sides:

**Example 6:** If the density of hydrogen is 0.090g/L and its rate of effusion is 5.93 times that of chlorine, what is the vapour density of chlorine?

**Solution:**

1. Set rates and weights:

Rate1 = H2 = 5.93

Rate2 = Cl2 = 1

The molecular weight of H2 = 2.016

The molecular weight of Cl2 = x.

2) By Graham's Law:

.93

3) Determine gas density using;

**Example 8:** How much faster does hydrogen escape through a porous container than sulphur (IV) oxide?

**Solution:**

1. Rates & weights:

rate1 = H2 = x

rate2 = SO2 = 1

The molecular weight of H2 = 2.0

The molecular weight of SO2 = 64.0

1. Graham's Law:
2. Substituting;

Hydrogen gas effuses 5.657 times as fast as SO2 gas does.

**Example 9:** Oxygen has an effusion rate of 10mL per second and a molecular mass of 32 amu. An unknown gas has an effusion rate of 5mL per second. Determine the molecular mass of the unknown gas.

**Solution**

**Example 10:** The rates of diffusion of an unknown gas and chlorine gas are in the ratio of 6:5. Assuming the density of chlorine to be 36, calculate the molecular weight of the unknown gas.

**Solution**

Square both sides;

Therefore, the relative density of gas B = 25. Hence, the molecular weight = 25 × 2 = 50 (Relative vapour density = molecular masse/2).

**Example 11:** 180mL of a hydrocarbon diffuses through a porous membrane in 15 minutes while 120mL of SO2 under the same identical conditions diffuses in 20 minutes. Calculate the molecular weight of the hydrocarbon.

**Solution**

Rate of diffusion of hydrocarbon = R1.

Similarly, rate of diffusion of SO2 = R2.

Molecular mass (MM2) of SO2 = 32 + 16 × 2 = 64 gmol-1

Square both sides:

The hydrocarbon is methane, CH4 = 12 +1 × 4 = 16

**Example 12:** 200cm3 of oxygen diffuses through a porous plug in 50 seconds. How long will 80cm3 of methane take to diffuse through the same porous plug under the same conditions?

**Solution**

Ro = Rate of diffusion of oxygen = volume/time = 200/50 = 4cms-1

Rm = Rate of diffusion of methane = ?

Mo = Molar mass of oxygen = 32g

Mm = Molar mass of methane = 16g

Now; Ro/Rm = √Mm/√Mo

4/Rm = √16/√32

Rm√16 = 4√32

Rm = 4√32/√16

Rm = 4 × 5.65685/4

Rm = 5.65

Recall: Rate = volume/time

5.65 = 80cm3/time

Time = 80/5.65

Time = 14seconds

**Example 13:** If sulphur (IV) oxide and methane are released simultaneously at the opposite ends of an arrow tube, what is the ratio of their rates of diffusion?

**Solution**

Molar mass of SO2 = (32 + 16 × 2) = 64g Vapour density of SO2 = Molar mass/2 = 64/2 = 32

Molar mass of CH4 = (12 + 1 × 4) = 16g vapour density of CH4 = Molar mass/2 = 16/2 = 8

Rs/Rm = √dm/√ds

Where,

Rs = rate of diffusion of SO2, Rm = rate of diffusion of methane, dm = density of methane and ds =density of SO2.

Rs/Rm = √dm/√ds

Rs/Rm = √8/√32

Rs/Rm = √1/√4

Rs/Rm = 1/2

Rate of SO2: CH4 = 1 : 2

**DENSITY**

In science, densities of solids and liquids are expressed in grams per cubic centimeter or grams per milliliter whereas the densities of gases are expressed in grams per litre.

***Density () is defined as the ratio of mass of a substance to its volume.***

**Example 1:** A clean dry 10.0 ml graduated cylinder weighs 37.6g empty; it weighs 53.2g when filled to the 7.4ml mark with an unknown liquid. Calculate the density of the liquid.

**Solution**

Weight of cylinder = 37.6g

Weight of cylinder and unknown liquid = 53.2g

Weight of the liquid = 53.2g - 37.6g = 15.6g

The volume of the liquid = 7.4ml

**Example 2:** A clean dry volumetric flask, known to contain exactly 10.0ml when it is properly filled to the mark, weighs 12.754g when empty. When filled to the mark with a liquid at 23°C, it weighs 33.671g. Calculate the density of the liquid.

**Solution**

Weight of volumetric flask = 12.754g

Weight of volumetric flask + liquid = 33.671g

Weight of liquid = 33.671 - 12.754 = 20.917g

Volume of the liquid = 10.0ml

**Example 3:** A cylindrical rod weighing 45.0g is 2.0cm in diameter and 15.0cm in length. Find the density.

**Solution**

If a solid has a regular geometric form, the density may be calculated from its weight and volume.

Mass = 45.0g

Volume of cylinder = πrl

Remember; radius = diameter/2 = 2cm/2 = 1.0cm

Volume of cylinder = (3.14)(1.0cm)2(15.0) = 47.1cm3

Example 4: A 5.7 g sample of metal pellet is put into a graduated cylinder that contains 5.0ml of water. After the pellets are added, the water level stand at 7.7mL. What is the density of the pellets?

**Solution**

Mass of pellets = 5.7g

Volume of pellets = Volume of water displaced = 7.7 - 5.0 = 2.7ml

Density of pellets = Mass/Volume

= 5.7/2.7

= 2.11

**Example 5:** What is the density of a solid cube that has a length of 2.0cm on each side and weighs 6.0 grams?

**Solution:**

Because the length of each side of this cube is 2.0cm, the volume will be the length × width × height. This is (2.0cm)(2.0cm)(2.0cm) = 8cm3. The mass is 6.0 grams.

Substitute into the equation;

The density of this solid is 0.75grams/cm3.

**TAKE AWAY**

1. Five liters of a gas exist at a pressure of 700 torr. If the pressure is increased to 1400, the volume is decreased. Determine the new volume. **Ans. 2.5 litres.**
2. What is the mass of an object that has a density of 13g/mL and a volume of 10mL? **Ans. 130g.**
3. The relative rates of diffusion of two gases A and B are found to be 0.3 and 0.2 respectively. If the density of A is 4, find the relative density of the gas B? **Ans. 9**
4. If 30cm3 of Oxygen diffuses through a porous plug in 7 seconds, how long will it take 60cm3 of chlorine to diffuse through the same plug? [O = 16, Cl = 35.4] **Ans. 21 seconds**

5. 350 cm3 of a gas diffuses through a porous pot in 10 min. What is its rate of diffusion? **[Answer: 0.6 cm3s−1]**

6. In an experiment to determine the rate of diffusion of a gas, a student discovered that 650 cm3 of the gas diffused in 136.3 s. What was the error in his result if the actual rate of diffusion of the gas is 4.71 cm3s−1? **[Answer: 0.10 s]**

7. 180 cm3 of a gas diffuses in 20 seconds. How long will it take 400 cm3 of oxygen to diffuse under the same conditions if the gas has a vapour density of 15? (O=16) **[Answer: 46 seconds]**

8. The rate of diffusion of methane is 3 times as fast as that of a gas x. What is the molar mass of x? (C = 12, H = 1) **[Answer: 144 g mol−1]**

9. A cylinder contains three gases, A, B and C whose amounts are 0.2 mol, 0.8mol and 1.2 mol respectively. Calculate the partial pressure of each gas if the gases occupy a volume of 2.5 dm3 at 1.1 atm and a volume of 2.5 dmlculate −1 mol−1) **[Answer: PA = 2 atm, PB = 6.8 atm, PC = 11atm]**

10. A mixture of gases contains four gases, A, B, C and D. The partial pressures of the gases are 2.5 atm, 0.8 atm, 1.9 atm and 2.7 atm respectively. Calculate the mole fraction of each gas. **[Answers: PA = 0.32,PB = 0.1, PC = 0.24, PD = 0.34]**

11. You are required to produce 500 cm3 of ammonia through the Haber process. What volumes of nitrogen and hydrogen would you require to obtain this result? The equation of reaction is

N2(g) + 3H2(g) ⇔ 2NH3(g).

**[Answer: N2 = 250 cm3, H2 = 750 cm3]**

12. What number of moles of oxygen would exert a pressure of 10 atm at 320 K in an 8.2 dm3 cylinder? [R = 0.082 atm dm3 mol-1 K-1] **Ans. = 3.13 moles**

13. 250 cm3 of carbon monoxide was sparked with 1050 cm3 of air containing21% of oxygen. Calculate the total volume of the residual gases after the resulting mixture is passed through alkaline pyrogallol? The equation of reaction is

2CO(g) + O2(g) (g)ti2(g).**[Answer: 830 cm3]**

14. 0.421 mol of a gas occupies a volume of 5.00 dm3 at 0.421 mol of a g pressure of the gas in Torr?(R = 62.4 L Torr K−1 mol−1) **[Answer: 1430 Torr]**

15. A gas occupies a volume of 30.0 dm3 at s.t.p. What volume in dm3 would it occupy at 91 and 380 mmHg? Ans. = 80 dm3

**CHAPTER 8: SOLUBILITY AND pH OF SOLUTIONS**

**SOLUBILITY**

**Solubility is the maximum amount of solute that would dissolve or be contained in 1.0 dm3 of solution at a given temperature.**

Solubility is measured either in grams per 100g of solvent (g/100g) or number of moles per 1L of the solution (mol/L). As an example, calculate the solubility of sodium nitrate, NaNO3, if 21.9g of the salt is dissolved in 25g of water. Based on this calculation, the final volume of the NaNO3 saturated solution is 55ml. The above question can be solved by following the steps below:

1. Calculate the molar mass of the dissolved compound as the sum of mass of all atoms in the molecule.

Molar mass of NaNO3 = 23 + 14 + 3 x 16 = 85gmol-1.

1. Divide the mass of the dissolved compound (21.9g) by its molar mass to calculate the number of moles. This would be: Number of moles (NaNO3) =
2. Divide the number of moles by the solution volume in liters or dm3 to calculate solubility in mole/L. In our example, the solution volume is 55mL or 0.055L. The solubility of NaNO3 = 0.258moles/0.055L = 4.69mole/L.

OR

Using the formula below:

Molar mass NaNO3 = 23 + 14 + 3 x 16 = 85 gmol-1.

**Example 2:**

If 1.11g of calcium chloride (CaCl2) were dissolved in 20.0g of distilled water at 20°C, calculate the solubility of the salt in moldm-3.[Ca= 40.0,Cl = 35.5]

**Solution**

Molar mass of CaCl2 =40 + (35.5 × 2) = 111gmol-1

Or

Molar mass of CaCl2 = 40 + (35.5 × 2) = 111gmol-1

20g of water at 20°Cdissolves 0.01 mol of CaCl2

**Example 3:**

Given that sodium chloride has a solubility of 36.3 at 30 and 39.0 at 100 and that of silver nitrate is 297.0 at 30and 952.0 at 100. Calculate the percentage of NaCl and AgNO3 deposited.

1. Deduce which of the two salts can be purified more efficiently by crystallization.

**Solution**

Mass of NaCl deposited = 39.0 - 36.3 = 2.7g

Mass of AgNO3 deposited = 952 – 297 = 655g

AgNO3 (Silver trioxonitrate (V)) will be more efficiently crystallized because it has a higher percentage deposited.

**Example 4:** At 25 °C, evaporation of a 100 cm3 solution of K2CO3to dryness gave 14 g of the salt. What is the solubility of K2CO3 at 25 °C? [K2CO3= 56]

**Solution**

1. Calculate the molar mass of the dissolved compound as the sum of mass of all atoms in the molecule. Molar mass K2CO3 = 39× 2 + 12× 1 + 16 x 3 = 138 g/mol.

2. Divide the mass of the dissolved compound (14g) by its molar mass to calculate the number of moles. This would be: Number of moles (K2CO3) =14g /138g/mole = 0.005 moles.

3. Divide the number of moles by the solution volume in liters or dm3 to calculate solubility in mole/L or mol/dm3. In our example, the solution volume is 100 cm3or 0.1dm3. The solubility of K2CO3 = 0.005 mol/0.1dm3= 1.01 moldm-3.

Alternatively;

**Example 5:** Water was added to 120g of a salt MCl2 to produce 60 cm3 of a saturated solution at 298K. The solubility of the salt at that temperature was 8.0   
. Calculate the mass of the salt which remains undissolved. [M = 24, Cl = 35.5].

**Solution**

=

**SOLUBILITYPRODUCT**

Solubility product constant, Ksp is simply equilibrium constant. Hence, Ksp is the equilibrium constant for a solid substance dissolving in an aqueous solution (It represents the dynamic equilibrium between a solid substance and its dissociated ions in a saturated solution).

**It equals the product of the equilibrium concentrations of the ions in the compound, each concentration raised to a power equal to the number of such ions in the formula of the compound. Ksp is temperature dependent.**

The Ksp for the salt, CaF2:

CaF2(s)Ca2+(aq)+2F-(aq)

Ksp = [Ca2+][F-]2

Where [Ca2+] is the solubility or concentration of calcium ion in moldm-3, [F-]2 is the solubility or concentration of fluoride ion in moldm-3.

The abbreviation Ksp stands for solubility product constant. It is called a product because the negative and positive ionconcentrationsaremultipliedtogethertodeterminethevalueoftheconstant.

**What factors affect the value of Ksp?**

✓ The common-ion effect (the presence of a common ion lowers the value of Ksp).

✓ The diverse-ion effect (if the ions of the solutes are uncommon, the value of Ksp will be high).

**Example 1:**

Write the solubility product expression for the saturated aqueous solution of Bi2S3.

**Solution**

Bi2S3(s)  2Bi3+(aq) + 3S2-(aq)

Ksp = [Bi3+]2[S2-]3

**Example 2:** Calculate the solubility product of nickel hydroxide (Ni(OH)2), if the solubility is 1.86×10-6 moldm-3.

**Solution**

Ni(OH)2 ⇌ Ni2++ 2OH-

[Ni2+] = 1.86×10-6 moldm-3

[OH-] = 2 × 1.86 × 10-6moldm-3

= 3.72 × 10-6 moldm-3

Ksp = [Ni2+][OH-]2

= (1.86×10-6)(3.72×10-6)2

= 2.574 × 10-17mol-3dm-6

**Example 3:**

A litre of a saturated solution at 25°C with calcium oxalate, CaC2O4 is evaporated to dryness, giving a 0.0061g residue of CaC2O4. Calculate the solubility product constant for this salt at 25°C.

**Solution**

First, find the concentration of CaC2O4 in moldm-3:

Molar mass of CaC2O4 = 40 + 12 × 2 + 16 × 4 = 128g/mol

CaC2O4(s) <==>Ca2+(aq)+C2O42-(aq)

Ksp = [Ca2+][C2O42-]

= (4.8×10-5)(4.8×10-5 )

Ksp = 2.3×10-9

**Example 4:**

Silver ion may be recovered from used photographic fixing Solution by precipitating it as silver chloride. The solubility of silver chloride is 1.9×10-3g/L. Calculate Ksp.

**Solution**

Convert the concentration of silver chloride from g/L (g/dm3) to mol/L(mol/dm3):

Molar mass of AgCl = 108 + 35.5 = 143.5g/mol

AgCl(s)<==>Ag+(aq)+Cl-(aq)

Ksp = [Ag+][Cl-]

= (1.324×10-5)(1.324×10-5)

Ksp = 1.7529×10-10

= 1.8×10-10

**pH AND pOH OF SOLUTIONS**

**INTRODUCTION**

pH of a solution is a measure of the negative logarithm of hydrogen or hydroxonium ions concentration.

pH = -log[H+] or pH = -log[H3O+]

**N/B: The brackets [] means concentration or Molarity.**

pOH of a solution is a measure of the negative logarithm of hydroxyl ions concentration.

pOH = - log[OH-]

**A neutral Solution, whose hydrogen ion concentration [H+] or hydroxonium ions concentration [H3O+] at 25°C is 1.0 × 10-7 moldm-3, has a pH of 7.00. For an acid solution, the hydrogen ion concentration is greater than 1.0 × 10-7 moldm-3, so the pH is less than 7.00. Similarly, a basic Solution has a pH greater than 7.00.**

[H3O+][OH-] = 1.0 × 10-14

Taking the logarithm of both sides of the equation, you get; (-log[H3O+]) + (-log[OH-]) = 14.00

Hence, pH + pOH = 14.00

To find the concentration of the solution (H+ or OH-), use:

**Example 1:** If [H+] = 1.00 X 10-10M, what is the pH?

**Solution**

pH = -log(1X10-10)

= 10

pH=10.0 × 1 = 10.0

**Example 2:** If [OH-] = 1.80 X 10-5M, what is the pOH?

**Solution**

pOH =-log 1.80 X 10-5

pOH = 4.74

**Example 3:** What is the pH of the solution if the hydroxonium ions concentration is 2.5 × 10-4 moldm-3?

**Solution**

pH = -log[H3O+]

= -log[2.5 × 10-4]

= 3.6

**Calculating the pOH or pH when you know the other One:**

Since acids and bases are opposites, pH and pOH are opposites. If pH = 1, then pOH = 13. pH and pOH are on opposite ends, **pH + pOH = 14**

**Example 4:**  The pOH of the solution is 4.5. Calculate the pH of the solution.

**Solution**

pH + pOH = 14.00

pH + 4.5 = 14.00

pH = 14 - 4.5

pH = 9.5

Hence, at 25°C the solution will be basic since the pH value is greater than 7.00 i.e [OH-] >[H3O+]

**Example 5:** If pOH is 3.8. What is the hydroxyl ion concentration?

**Solution**

pOH = -log[OH-]

[OH-] = 10-pOH

= 10-3.8

[OH-] = 1.58 × 10-4 moldm-3

**Example 6:** What is the pOH of the solution if the hydroxonium ions concentration is 4.2 × 10-3 moldm-3?

**Solution**

pH = -log[H3O+]

= -log[4.2 × 10-3]

= 2.377

pH + pOH = 14.00

pOH = 14 - 2.377

= 11.623

**Example 7:** If [H3O+] is 7.1 × 10-2 M. Calculate the [OH-].

**Solution**

pH = -log[H3O+]

= -log [7.1 × 10-2]

= 1.149

pH + pOH = 14.00

pOH = 14 - 1.149

pOH = 12.851

[OH-] = 10-pOH

= 10-12.851

[OH-] = 1.41 × 10-13M

**Alternatively;**

H3O+][OH-] = 1.0 × 10-14

[OH-] = 1.41×10-13M

**There are several ways to test pH**

**✓** Blue litmus paper (red=acid), Red litmus paper (blue = basic)

✓ pH paper (multi-colored)

✓ pH meter (7 is neutral,<7 = acid, >7 = base).

**TAKE AWAY**

1. What is the [H+] of a solution that has an [OH-] = 4.20 x1 0-5?

2. What is the pH of a 1.70 x 10-5M solution of HCl? **Ans. pH = 4.77**

3. What is the pOH of a 0.320M solution of NaOH?

4. A solution has a pH of 6.40, what is the [H+]? 3.98 x 10-7 M?

5. WhatisthepOHofasolutionwithapHof4.50?

6. What is the [H+] of a 1.6x10-9M solution of NaOH?

7. Calculate the solubility of calcium fluoride in water from the solubility product constant of 3.4×10-11. **Ans. 2.0×10-4mol/dm3of CaF2.**

8. 65 g of potassium trioxonitrate (V) form a saturated solution with 100 cm3 of water at a certain temperature. Calculate the solubility of potassium trioxonitrate (V). [K = 39, N = 14, O = 16]. **Ans. 6.44 moldm-3**.

9. 3.06 g of a sample of potassium trioxonitrate (V) was required to make a saturated solution with 10 cm3 of water at 25 equal to the number of such ions in the formula of the compound. Kerature by s**Ans. 2.50 moldm-3**

10. 16.55g of lead (II) trioxonitrate (V) was dissolved in 100g of distilled water at 20°C. Calculate the solubility of the solute. [Pb = 207, N = 14, O = 16]. **Ans. 0.5 moldm-3**.

**CHAPTER 9: ACID-BASE REACTIONS**

**TITRATION**

**Titration** is defined as the slow addition of one solution of a known concentration (called a titrant) from the burette to a known volume of another solution of unknown concentration in a conical flask until the reaction reaches neutralization, which is often indicated by a color change. A titration is a method of analysis that will allow you to determine the precise endpoint of a reaction and therefore the precise quantity of reactant in the titration flask.

is a dye that changes colour when pH changes.

is when the indicator changes colour during a titration.

is when the amount of acid and of base is just sufficient to cause complete consumption of both the acid and the base.

* At the equivalence point neither the acid nor the base is in excess.
* At the equivalence point neither the acid nor the base is the limiting reagent.
* The pH of the solution at the equivalence point depends on the relative strength of the acid and strength of the base used in the titration.
* The nature of this aqueous salt solution determines the pH of the resultant solution.

**In general, we can apply the following generalization for aqueous solutions at 25olution determines the pH of the resultant solution.he acid and strength of the base used in**

* If acid is stronger than base, salt solution has a pH < 7 (acidic), use either methyl orange or phenolphthalein.
* If acid is weaker than base, salt solution has a pH > 7 (basic), use phenolphthalein.
* If acid strength is the same as base strength (strong acid against strong base), salt solution has a pH = 7 (neutral), use methyl orange.

**OPERATIONAL STEPS IN CARRYING OUT A TITRATION**

Many students make errors during titration due to improper use of volumetric apparatus. Tests of titrations are in fact test of correct use of the equipment.

**Rinsing the Pipette**

A pipette is used to measure and transfer a precise volume of liquid (measure out fixed volume of liquid very accurately). You rinse a pipette with the solution whose volume you are measuring. This ensures that the solution will not be diluted or contaminated.

1. Pour a sample of standard solution into a clean, dry beaker. Place the pipette tip in a beaker of distilled water.

2. Rinse the pipette by drawing several milliliters of solution from the beaker into it. Rotate and rock the pipette to coat the inner surface with solution. Discard the rinse. Rinse the pipette twice in this way. It is now ready to fill with standard solution.

**Filling the Pipette**

3. Place the tip of the pipette below the surface of the solution.

4. Use a pipette and pipette filler to measure 25.0cm3 of the base.

Note: Draw a bit more liquid than you need into the pipette. It is easier to reduce this volume than it is to add more solution to the pipette. The bottom of the meniscus must align exactly with the etched mark. You can prevent a “revents.e drop from clinging to the pipette tip by touching the tip to the inside of the conical flask.

***Never use your mouth instead of a pipette filler to draw a liquid into a pipette. The liquid could be corrosive or poisonous. As well, you would contaminate the glass stem.***

In case there is no pipette filler, pipette the base by sucking it above the mark. Close the upper end of the pipette with the dry index finger. The base is gradually allowed to run off by releasing the pressure of the finger while using the left hand to gradually rotate it until the bottom of the meniscus is on the mark at eye level. The base is then run into the conical flask by removing the finger from the top. If the base is sucked into the mouth, spit it out immediately and irrigate the mouth with plenty of water several times.

**Transferring the Solution**

5. Place the tip of the pipette against the inside glass wall of the flask. Let the solution drain slowly, by removing our finger from the stem.

6. After the solution drains, wait several seconds and then touch the tip to the inside wall of the flask to remove any drop on the end.

Note: You may notice a small amount of liquid remaining in the tip. The pipette was calibrated to retain this amount. Do not try to remove it.

**Adding the Indicator**

7. Add two or three drops of indicator to the flask and its contents. Do not add too much indicator. Using more does not make the colour change easier to see. Also, indicators are usually weak acids. Too much can change the amount of base need for neutralization. You are now ready to prepare the apparatus for the titration.

**Rinsing the Burette**

A burette is a graduated glass tube with a tap at one end. It is used to accurately measure the volume of liquid added during a titration experiment.

8. To rinse the burette, close the tap and add about 10mL of distilled water from a wash bottle.

9. Tip the burette to one side, and roll it gently back and forth so that the water comes in contact with all inner surfaces.

10. Hold the burette over a sink. Open the tap, and let the water drain out. While you do this, check that the tap does not leak. Make sure that it turns smoothly and easily.

11. Rinse the burette with 5mL to 10mL of the solution that will be measured. Remember to open the tap to rinse the lower portion of the burette. Rinse the burette twice, discarding the liquid each time. The rinsing is necessary in order not to decrease the concentration of the acid by the water left on the sides of the burette. If you are right-handed, the tap should be on your right as you face the burette. Use your left hand to operate the tap. Use your right hand to swirl the liquid in the Erlenmeyer flask. If you are left-handed, reverse this arrangement. Observe the level of solution in the burette so that your eye is level with the bottom of the meniscus.

**Filling the Burette**

12. Assemble a retort stand and burette clamp to hold the burette. Clamp the burette vertically firmed because if it is slant, the acid level will be tilted to one side. Thus, the correct reading becomes impossible. Place a funnel in the top of the burette.

13. With the tap closed, add solution until the liquid is above the zero mark. Remove the funnel immediately to prevent acid left in it from draining into the one already in the burette, thereby altering the readings. Carefully open the tap. Drain the liquid into a beaker until the bottom of the meniscus is at or below the zero mark.

14. Touch the tip of the burette against the beaker to remove any clinging drop. Check that the portion of the burette that is below the tap is filled with liquid and contains no air bubbles. This is necessary because the volume of acid in that portion adds to give whatever the burette reading is above. If it is not filled, the titre value will tend to be more than that actually used.

15. Record the initial burette reading in your notebook.

16. Replace the beaker with the Erlenmeyer conical flask that you prepared earlier. Place a sheet of white paper or white tile under the Erlenmeyer to help you see the indicator colour change that will occur near the endpoint. Near the endpoint, when you see the indicator change colour as liquid enters the flask from the burette, slow the addition of liquid. The endpoint can occur very quickly.

17. Slowly add the acid from the burette to the conical flask, swirling to mix. (The mixture may at first change colour, and then back again when swirled). As the end point is being approached, flashes of the colour change begin to show up.

18. Stop adding the acid when the end-point is reached (when the colour first permanently changes). Note the final volume reading and record on the answer booklet.

**What is a titre value? When the initial reading is subtracted from the final reading, the volume of acid obtained is called the titre value.**

19. Repeat the steps twice or thrice until results are repeatable (in close agreement). Ideally, these should lie within 0.10cm3 of each other.

**Washing a pipette or burette**

20. This should be washed with pure water (probably with a tap water rinse first) followed by some of the solution you are going to put in them. If you left any water in the burette or pipette, then the solution you put in them would be diluted by that water, and so is not the concentration you think it is.

**Reading the Burette**

21. A meniscus reader is a small white card with a thick black line on it. Hold the card behind the burette, with the black line just under the meniscus. Record the volume added from the burette to the nearest 0.01mL. A meniscus reader helps you read the volume of liquid more easily.

**MAKING SURE THAT TITRATION RESULT IS ACCURATE AND RELIABLE**

**Accuracy**

**✓** It is quite common to do a rough titration to start with, especially if you haven't had a lot of practice. You just add the liquid from the burette 1cm3 at a time until the indicator changes colour. For example, you might find that it hadn't changed colour after 24cm3, but had changed colour after 25cm3.

✓ You can now find an accurate result by quickly running in 24cm3 (which you know is safe), and then adding the rest very slowly-a drop at a time (or even a partial drop at a time) as you near the endpoint.

✓ So, if you are asked how to make your titration accurate, your answer is to add the liquid from the burette drop-wise as you get near the endpoint.

**Reliability**

✓ To check the reliability of your result, you repeat the titration several times until you have concordant titres. These are results which are within 0.10cm3 of each other. (The titre is the volume of liquid added to reach the end-point). One drop of the acid from the burette is equal to 0.05 cm3.

✓ When you average out your titration results, you should only use your concordant titres. If the rough titre value appears to be concordant with some of the later titre values, it can be used for averaging.

**Sample of Titration Results**

**Indicator used: Methyl orange (based on the instructions). Volume of pipette: 25.0cm3 or 20.0cm3**

| Burette Reading | Rough (cm3) | 1st titre (cm3) | 2nd titre (cm3) | 3rd titre (cm3) |
| --- | --- | --- | --- | --- |
| Final reading | 24.50 | 23.80 | 47.50 | 23.90 |
| Initial reading | 0.00 | 0.00 | 23.80 | 0.00 |
| Volume of acid used | 24.50 | 23.80 | 23.70 | 23.90 |

**Note: Always take your burette readings to 2 decimal places and the difference in the titre values should not be more than 0.20 cm3.**

Calculate the mean titre. In the calculation, ignore any results that are not in close agreement (24.50cm3 in the table above) or results that are not concordant.

**Check your progress**

1. Explain why a pipette is used to measure the base, rather than a measuring cylinder.

**Ans. A pipette is more precise than a measuring cylinder. Adding slightly different volumes of alkali to the flask will result in a systematic error** (an error in measurement which differs from the true value by the same amount each time. This could be due to the equipment, how the experiment is carried out or the environment)**. The pipette allows the same volume of base to be added each time, helping to make the results repeatable.**

1. Explain importance of a suitable indicator in obtaining accurate results.

**Ans. The indicator must change colour sharply when the solution in the flask is neutralized. This means the volume of acid measured is very close to the true value. The white tile makes it easier to see the colour change. Misjudging the colour change could result in a random error.**

1. Describe two steps needed to obtain accurate results during titration.

**Ans. Take the readings from the bottom of the meniscus. Near to the end-point, rinse the inside of the flask with distilled water and add the acid drop by drop.**

**CALCULATING THE CONCENTRATION OF ACID IN THE STOCK SOLUTION**

To calculate the concentration or Molarity of a stock solution based on the information from the reagent bottle label, use the formula below:

Supposed, a HCl reagent bottle label carries the following information among others:

Specific gravity = 1.18g/cm3

Assay (% purity) = 36%

Molar mass = 36.5g/mol

Calculate the concentration of HCl in mol/dm3.

**Solution**

**NOTE: Specific gravity × 1000 = mass of 1 dm3 of solution.**

**CALCULATING MOLARITY FROM MASS AND VOLUME**

**Example 1:** A sample of NaNO3 weighing 0.38 g is placed in a 50.0 mL volumetric flask. The flask is then filled with water to the mark on the neck, dissolving the solid. What is the molarity of the resulting solution?

**Solution**

To find the molarity, you need the moles of solute, NaNO3. Therefore, you first convert grams of NaNO3 to moles. The molarity equals the moles of solute divided by the volume of solution in dm3.

Molar mass of NaNO3 = 23 + 14 + 16 × 3 = 85gmol-1

Convert the 50 cm3 to dm3 by dividing it by 1000 cm3:

Now that we have obtained amount in mole and the volume in dm3, we can get the molarity using the formula below:

**Example 2:** A sample of sodium chloride, NaCl, weighing 0.0678 g is placed in a 25.0cm3volumetric flask. Enough water is added to dissolve the NaCl, and then the flask is filled to the mark with water and carefully shaken to mix the contents. What is the molarity of the resulting solution?

**Solution**

To find the molarity, you need the moles of solute, NaCl. Therefore, you first convert grams NaCl to moles. The molarity equals the moles of solute divided by the volume of solution in dm3.

Molar mass of NaCl = 23 + 35.5 = 58.5gmol-1

Convert the 25.0cm3 to dm3 by dividing it by1000cm3:

Now that we have obtained amount in mole and the volume in dm3, we can get the Molarity using the formula below:

**Example 3:** A flask contains a solution with an unknown amount of HCl. This solution is titrated with 0.207 M NaOH. It takes 4.47 cm3 NaOH to complete the reaction. What is the mass of the HCl?

**Solution**

**Before you perform the titration calculations, always write down the balanced chemical equation.**

HCl + NaOH => NaCl + H2O

From the balanced chemical equation, 1 mole of NaOH reacted with 1 mole of HCl.

Mole of NaOH = Concentration × volume

Since 1 mole of NaOH reacted with 1 mole of HCl, 0.00092529mole of NaOH will react with 0.00092529 mole of HCl.

Mass of HCl = Amount × Molar mass

Molar mass of HCl = 1 + 35.5 = 36.5 gmol-1

Mass of HCl = 0.00092529 × 36.5 = 0.03377g

Mass of HCl = 0.0338g

**Example 4:** A 5.00-g sample of vinegar is titrated with 0.108 M NaOH. If the vinegar requires 39.1cm3 of the NaOH solution for complete reaction, what is the mass percentage of acetic acid, HC2H3O2, in the vinegar?

The reaction is HC2H3O2(aq) + NaOH(aq) => NaC2H3O2(aq) + H2O(l)

**Solution**

Convert the volume of NaOH solution (0.0391 dm3 i.e 39.1/1000) to moles NaOH (from the molarity of NaOH):

= 0.108 × 0.0391 = 0.0042228 mole of NaOH.

Then convert moles NaOH to moles HC2H3O2 (from the chemical equation):

1 mole of NaOH = 1 mole of HC2H3O2

0.0042228 mole of NaOH will be 0.0042228 mole of HC2H3O2

Finally, convert moles of HC2H3O2 to grams HC2H3O2 :

Molar mass of HC2H3O2 = 1 × 1 + 2 × 12 + × 1+ 2 × 16 = 60.0 gmol-1

Reacting mass = 0.0042228 × 60.0

Reacting mass = 0.253368 g of HC2H3O2

The mass percentage of acetic acid in the vinegar can now be calculated:

**CALCULATING THE CONCENTRATION OF ACIDS AND BASES (in gdm-3 or moldm-3)**

* **Example 1**: A is a solution of HCl acid. B is a solution containing 4g of NaOH per dm3 solution. Given that 25.0 cm3portions of B neutralizes 29.50 cm3 of A. Calculate the concentration of A in moldm-3.

**Solution**

NaOH(aq) + HCl(aq) => NaCl(aq) + H2O(aq)

1 : 1

To find the concentration of A in moldm-3, we have to get the concentration of B in moldm-3 and not in gdm-3 :

The molar mass of B (NaOH) = 23 + 16 + 1 = 40

We then calculate concentration of A in moldm-3 using the formula below:

CA = concentration of acid in moldm-3 =?, CB = concentration of base in moldm-3 = 0.100 moldm-3, VA = Average volume of acid = 29.50cm3, VB = volume of base = 25.00cm3, nA = number of moles of acid in the balanced chemical equation = 1 mole, nB = number of moles of base in the balanced chemical equation = 1 mole. Substituting the values;

CA × 29.50 × 1 = 0.100 × 25 × 1

CA = 0.0847 moldm-3 (3 - significant figures).

**Example 2:** A is a solution of dilute trioxonitrate (V) acid of unknown concentration. B is a solution containing 5.34g of sodium trioxocarbonate (IV) per 1 dm3 Solution.

**Procedure:** Put A in the burette and titrate against 25 cm3 portions of B using methyl orange as indicator. Tabulate your burette readings. From your results obtained, Calculate:

a) The average volume of A used

b) Concentration of A in moldm-3

c) Concentration of A in gdm-3

Equation for the Reaction:

2HNO3(aq) + Na2CO3(aq) =>2NaNO3(aq) + H2O(l) + CO2(g)

2 : 1

Volume of pipette = 25.00 cm3

Indicator used = Methyl orange

Colour of indicator in Na2CO3 = Yellow

Colour change at endpoint = orange

**Titration Results**

| Burette Reading | Rough(cm3) | 1st titre(cm3) | 2nd titre(cm3) | 3rd titre(cm3) |
| --- | --- | --- | --- | --- |
| Final reading | 24.50 | 34.50 | 26.50 | 24.60 |
| Initial reading | 0.00 | 10.00 | 2.00 | 0.00 |
| Volume of acid used | 24.50 | 24.50 | 24.50 | 24.60 |

b) Concentration of A in moldm-3 (molar Concentration):

But before we apply the above formula, we have to find the concentration of B in moldm-3.

Mass Concentration = molar Concentration × Molar mass

The molar mass of Na2CO3 = 23 × 2 + 12 × 1 + 16 × 3 = 106gmol-1.

Mass Concentration = molar Concentration × Molar mass

5.34 = molar Concentration × 106

Having obtain the Concentration of B in moldm-3, we can now find the concentration of A in moldm-3, using

CA × 24.50 × 1 = 2 × 0.0504 × 25

CA = 0.0540 × 25 × 2/24.50

CA = 0.103 moldm-3 or 0.103 M. Where M = Molarity

c) The Concentration of A in gdm-3:

Mass concentration (g/dm3) = molar Concentration × molar mass

The molar mass of HNO3 = 1 + 14 + 16 × 3 = 63g/mol

Mass concentration = 0.103 × 63 = 6.49 g/dm3.

**Example 3:** F is 7.30g hydrochloric acid per dm3 Solution.

G is 10.6g of X2CO3 per dm3 Solution. Put F into burette and titrate against 25 cm3 portions of G using methyl orange as indicator. Assuming that the results of the titration were as follows:

| Burette Reading | Rough (cm3) | 1sttitre (cm3) | 2ndtitre (cm3) |
| --- | --- | --- | --- |
| Final reading | 26.40 | 29.30 | 26.20 |
| Initial reading | 0.00 | 03.00 | 0.00 |
| Volume of acid used | 26.40 | 26.20 | 26.20 |

From the result above, calculate:

a) Concentration of F in moldm-3

b) Concentration of G in moldm-3

c) Molar mass of G

d) The relative atomic mass of X in X2CO3

e) The mass of X2CO3 needed to prepare 1dm3 of

**Solution**

a) Find the Concentration of F in moldm-3:

b) Find the concentration of G in moldm-3 using;

CF = 0.200 moldm-3, VF = 26.2cm3, nF = 2moles, nG = 1 mole, VG = 25.0cm3

0.200 × 26.20 × 1 = CG × 25 × 2

CG = 0.105 moldm-3

c) Find the molar mass of G:

d) Find the relative atomic mass of X in X2CO3:

Molar mass of X2CO3 = 101

2x + 12 + 16 × 3 = 101

2x + 60 = 101

2x = 101 - 60

2x = 41

X = 41/2

X = 20.5

R.A.M of X = 20.5

e) Find the mass of X2CO3 needed to prepare 1 dm3 of 1.5 moldm-3

= 1.5

Number of moles of X2CO3 = 1.5moles

Mass in grams = number of moles × molar mass

Mass in grams of X2CO3 = 1.5 × 101

Mass in grams of X2CO3 = 151.5g

**Example 4**: A is a solution of hydrochloric acid. B is a solution containing 2.45g of anhydrous sodium trioxocarbonate (IV) in 250 cm3 of Solution. Put A into the burette and titrate it against 20 cm3 or 25cm3 portions of B using methyl orange as indicator. Repeat the exercise to obtain consistent titres. Tabulate your burette readings and calculate the average volume of A used. The equation for the reaction involved in the titration is:

Na2CO3(aq) + 2HCl(aq) => 2NaCl(aq) + H2O(l)+ CO2(g)

b) From your results and the information provided, Calculate the:

I) Concentration of B in moldm-3

II) Concentration of A in moldm-3

III) Concentration of A in gdm-3

IV) Volume of the gas evolved in the reaction at s.t.p.

[H = 1, C = 12, O = 16, Na = 23, Cl = 35.5, Molar volume = 22.4 dm3mol-3]

**Solution**

Titration Results

**Volume of pipette used: 25.0 cm3, Indicator used: Methyl orange**

| Burette Reading | Rough (cm3) | 1st titre (cm3) | 2nd titre (cm3) |
| --- | --- | --- | --- |
| Final reading | 23.90 | 23.80 | 23.80 |
| Initial reading | 0.00 | 0.00 | 0.00 |
| Volume of acid used | 23.90 | 23.80 | 23.80 |

bi) Concentration of B in moldm-3: First, convert the volume in cm3 to dm3;

1000 cm3 = 1dm3

250 cm3 = ?

Cross multiplying, we get 250 cm3 × 1 dm3/1000cm3 = 0.25dm3.

To find the Concentration of B inmoldm-3, we use the formula below:

Mass Concentration (g/dm3) = Molar Concentration (moldm-3) × Molar mass (g/mol)

Mass Concentration of B = 2.45g/0.25dm3 = 9.8 g/dm3

Molar mass of B = 23 × 2 + 12 + 16 × 3 = 106 g/mol

bii) Concentration of A in moldm-3:

Na2CO3(aq) + 2HCl(aq) =>2NaCl(aq) + H2O(l) + CO2(g)

1 : 2

CA = ? VA = 23.80 cm3, CB = 0.092 moldm-3, VB = 25.00cm3, nA = 2 moles, nB = 1 mole.

Substituting the values, we get;

biii) Concentration of A in gdm-3:

Mass Concentration (g/dm3) = Molar Concentration (moldm-3) × Molar mass (g/mol)

Molar mass of A (HCl) = 1 + 35.5 = 36.5gmol-1

Mass Concentration (g/dm3) = 0.193 × 36.5 = 7.045 gmol-1.

biv) Volume of the gas evolved in the reaction at s.t.p.

Na2CO3(aq) + 2HCl(aq) => 2NaCl(aq) + H2O(l) + CO2(g)

From the balanced chemical equation;

1 mole of Na2CO3 = 1 mole of CO2

0.092 mole of Na2CO3 will contain 0.092 mole of CO2

1 mole of CO2 = 22.4 dm3 at s.t.p

0.092 mole of CO2 = ?

Cross multiplying, we get;

**Example 5:** A is a solution containing 15.8 g dm-3 of Na2S2O3. B was obtained by dissolving 9.0 g of an impure sample of I2 in aqueous KI and the solution made up to 1 dm3.

Put A into the burette. Pipette 20.0 cm3 or 25.0 cm3 of B. Use starch solution as indicator. Repeat the titration to obtain concordant titre values. Tabulate your results and calculate the average volume of A used.

The equation for the reaction involved in the titration is:

I2 + 2S2O32- → 2I- + S4O62-

From your results and information provided, calculate the:

1. Concentration of A in moldm-3;

2. Concentration of I2 in B in mol dm-3;

3. Percentage by mass of I2 in the sample.

4. Give reasons why the starch indicator was not added to the titration mixture at the beginning of the titration.

[ O = 16.0, Na = 23.0, S = 32.0, I = 127.0 ]

**Solution**

**Indicator used: Starch, volume of pipette used: 25.00cm3**

| Burette Reading | Rough (cm3) | 1sttitre (cm3) | 2nd titre (cm3) |
| --- | --- | --- | --- |
| Final reading | 24.90 | 24.50 | 24.50 |
| Initial reading | 0.00 | 0.00 | 0.00 |
| Volume of acid used | 24.90 | 24.50 | 24.50 |

Molar mass of Na2S2O3 = (2 × 23) + (2 × 32) + (3 × 16) = 158 g/mol

Concentration of Na2S2O3 in moldm-3 (CA) = 0.100 moldm-3

I2 + 2S2O32- → 2I- + S4O62-

1 : 2

3. Mass of I2 in sample = mole of I2 × molar mass of I2

Molar mass of I2 = 2 × 127 = 254g/mol

Mass of I2 = 0.049 × 254

= 12.446 g

4. To obtain an accurate endpoint or to prevent the formation of a complex

**Example 6:**  A is 0.100 mol dm-3 HNO3.B is a solution containing 2.50 g of a mixture of Na2CO3 and Na2SO4 in 25.0cm3 of solution.

(a) Put A into the burette and titrate it against 20.0cm3 or 25.0cm3 portion of B using methyl orange as indicator. Repeat the exercise to obtain consistent titres values. Tabulate your burette readings and calculate the average volume of A used.

The equation of reaction is involved in the titration is:

1. Na2CO3(aq) + 2HCl(aq) => 2NaCl(aq) + H2O(1) + CO2(g)

(b) From your results and the information provided, calculate the:

(i) Concentration of B in moldm-3;

(II Concentration of Na2CO3 B in gdm-3;

(iii) Percentage of Na2CO3 in the mixture. [Na2CO3 = 106]

Assuming the average volume of acid used is 24.80 cm3.

**Solution**

(b) (i) Conc. of B in moldm-3

(ii) Conc. of Na2CO3 in B gdm-3:

Mass Conc. = Molar Conc. × Molar mass.

Mass Conc.= 0.0496 × 106 = 5.2576 ~ 5.3gdm-3

(iii) Percentage of Na2CO3 in the mixture

1000cm3 of B contains 5.3gdm-3

250cm3 of B contains = ?

Cross multiplying we have;

**Example 7:**  A is a solution of 0.050 moldm-3 H2C204. B is a solution of KMnO4 (potassium tetraoxomanganate (VII), of unknown concentration.

A) Put B into the burette. Pipette 20.0cm3 or 25.0cm3 of A into a conical flask and add about 10.0cm3 of dilute H2SO4. Heat the mixture to about 40oC-50oC and titrate it while still hot with B. Repeat the titration to obtain consistent titre values. Tabulate your results and calculate the average volume of B used.

Equation for the Reaction:

2MnO4- + 5C2O5 2- + 16H+=> 2Mn2+ + 8H2O + 10CO2

From your results and information provided, calculate:

I. Concentration of MnO4- in B in moldm-3

II. Concentration of MnO4-in B in gdm-3

III. Volume of CO2 evolved at s.t.p when 25.00cm3 of H2C2O4 reacted completely.

[O = 16, K= 39, Mn. = 55, Molar volume = 22.4 dm3 mol-1]

**Solution**

**Titration Results**

| Burette Reading | Rough(cm3) | 1st titre (cm3) | 2nd titre (cm3) |
| --- | --- | --- | --- |
| Final reading | 23.70 | 23.70 | 23.70 |
| Initial reading | 0.00 | 0.00 | 0.00 |
| Volume of acid used | 23.70 | 23.70 | 23.70 |

II. Concentration of MnO4- in B in gdm-3 (mass Concentration) is calculated using the formula below:

Mass Concentration = Molar Concentration × Molar mass

Molar mass of KMnO4 = 39 + 55 + 16 × 4 = 158.0gmol-1

Mass Concentration = 0.0211 × 158 = 3.3338gdm-3

Mass Concentration = 3.33 gdm-3

III. Volume of CO2 evolved:

Amount of H2C2O4 used = Concentration (moldm-3) × Volume (dm3)

We have to convert 25.0m3 to dm3 by dividing it by 1000(25/1000 = 0.025)

= 0.05 × 0.025 = 0.00125mol

From the balanced chemical equation:

5 moles of C2O4-2 = 10 moles of CO2

0.00125 mole of C2O4 -2 = ?

Cross multiplying we get;

At s.t.p:

1 mole of CO2 = 22.4dm3

0.00250 mole = ?

Cross multiplying we get;

**Example 8:** What volume of 0.1 M KMnO4 solution is required to oxidize 25.0 mL of 0.20 M FeSO4 in acidic solution?

**Solution**

The redox equation is shown below:

**Example 9:** E is a solution of sodium hydroxide containing 4.0g per dm3 Solution. H is a solution of mineral acid containing 0.06 moles per dm3 Solution.

**Procedure:** Put H in a burette and titrate it against 25 cm3 portions of E using methyl orange as indicator. Tabulate your burette readings and find the average volume of the acid used. From your results, calculate:

a) The number of moles of acid H in the average titre.

b) The number of moles of NaOH in the volume of E pipette.

c) The mole ratio of acid to base in the reaction.

d) Suggest the type of acid

e) Write the equation for the reaction.

**Solution**

**Titration Results**

Indicator used: Methyl orange. Volume of pipette: 25.0cm3

| Burette Reading | Rough(cm3) | 1st titre(cm3) | 2nd titre(cm3) |
| --- | --- | --- | --- |
| Final reading | 20.60 | 20.50 | 22.50 |
| Initial reading | 0.00 | 0.00 | 02.00 |
| Volume of acid used | 20.60 | 20.50 | 20.50 |

a) Find the number of moles of mineral acid in average titre using the formula below:

Mole = Volume (dm3) × Molar Concentration (moldm-3)

Convert 20.50 cm3 to dm3:

1000 cm3 = 1dm3

20.50 cm3 = ? Cross multiplying, we get;

20.50 cm3 × 1 dm3 / 1000 cm3 = 0.0205 dm3

Number of moles of H = 0.0205 × 0.06 Number of moles of H = 0.00123 mole

b) Find the number of moles of NaOH in the volume of E using the formula below:

Mole = Volume (dm3) × Molar Concentration (moldm-3)

Convert 25 cm3 to dm3:

1000 cm3 = 1 dm3

25 cm3 =?

Cross multiplying, we get;

25 cm3 × 1 dm3 /1000 cm3 = 0.025 dm3

Molar mass of NaOH = 23 + 16 + 1 = 40.0 gmol-1

Number of moles of H = 0.025 × 0.100

Number of moles of H = 0.00250 mole

c) Find the mole ratio of acid to base in the reaction:

This means that the mole ratio of base to acid is:

d) The mineral acid H could be a dibasic acid, that is, it has two hydrogen atoms in its molecule. E.g H2SO4, hence, two moles of NaOH is required to neutralize one mole of the acid.

e) Equation for the Reaction:

2NaOH(aq) + H2SO4(aq) => Na2SO4**(aq)** + 2H2O(l)

**Example 9:**

In a titration experiment, 22.50 cm3 of an acid solution A containing 10.6 g of NaHSO4 per dm3 reacted with 25.0 cm3 of solution B containing X g of NaOH per dm3. The equation for the reaction is:

NaHSO4(aq) + NaOH(aq) Na2SO4(aq) + H2O(l)

(a) From the information given above, calculate the:

(i) Concentration of A in moldm– 3;

(ii) Concentration of B in moldm– 3;

(iii) Value of X;

(iv) Mass of Na2SO4 formed during the reaction. [ H = 1.00; 0 = 16.0; Na = 23.0; S = 32.0]

(b) (i) Name a suitable indicator for the titration experiment.

(ii) State the apparatus used to measure the volume of solution: I. A; II B.

**Solution**

(a) (i) Mass Conc. = Molar Conc. × Molar mass

Molar mass NaHSO4 = [23.0 + 1.00 + 32.0 + 64.0] = 120 gmol-1 (No score for wrong Unit)

(ii) Reaction: **NaHSO4(aq) + NaOH(aq) Na2SO4(aq) + H2O(l)**

iii) Mass Conc. = Molar Conc. × Molar mass

Molar mass of NaOH = (23.0 + 16.0 + 1.00) = 40 gmol-1 (No score for wrong Unit)

Mass Conc. of NaOH gdm-3 = 0.0795 × 40 = 3.18gdm-3

Value of X = 3.18

(iv)

Amount of NaOH in 25.0cm3 of B:

Amount = Volume (dm3) × Conc.(moldm-3)

= (25.0 x 0.0795)/1000 = = 0.00199 mol

**We divide by 1000 cm3 to convert 25.0 cm3 to dm3 (0.025 dm3). Recall, 1 dm3 is equivalent to 1dm3.**

From the reaction equation;

1 mol NaOH gives 1 mol Na2SO4

0.00199 mol NaOH = 0.00199 mol Na2SO4

Molar mass of Na2SO4 = (2 x 23.0) + 32.0 + (4 x 16.0)= 142gmol-1

Reacting mass = Molar mass × Amount

Mass of Na2SO4 = 142 x 0.00199 = 0.283g

**Alternative Method to (iv); Mass of NaOH in 25.0cm3 of 0.0795moldm-3**

Where, m = reacting mass in grams, M = Molar mass in gmol-1, C = Concentration in moldm-3 and V = volume in dm3.

Molar Mass of Na2SO4 = 142gmol-1

From the balanced equation:

40g NaOH = 142g of Na2SO4

0.0795g NaOH = ?g Na2SO4

Cross multiplying we get;

(b) (i) Phenolphthalein / methyl orange

(ii) I - Burette

II - Pipette

**WATER OF CRYSTALLIZATION**

Water of crystallization is the water molecules bonded to the crystals of a salt. The idea behind it is that many bases also exist as crystalline solids (e.g Na2CO3), so therefore, when titrating them with a standard Solution of an acid (e.g known concentration) allows you to work out the moles of the base and by difference, the moles of water.

**NOTE: If you dissolve a known mass of hydrated solid base in water, you can work out the moles of anhydrous base, together with a known number of moles of water (X).**

**.**

**Example 1:** 6.43g of a sample of hydrated sodium trioxocarbonate (IV) was dissolved in 250 cm3 of water. A 25 cm3 of this solution was found to need 19.7cm3 of 0.230 moldm-3 H2SO4 for neutralization.

Na2CO3 + H2SO4 => Na2SO4 + CO2 + H2O

Calculate the relative molecular mass of the hydrated sodium trioxocarbonate IV and hence, the value of X.

**Solution**

Equation of Reaction:

Na2CO3 + H2SO4 =>Na2SO4 + CO2 + H2O

1 mole : 1 mole

Convert 6.43g of hydrated sodium trioxocarbonate in 250cm3 to g per dm3.

1000cm3= 1dm3

250cm3 = ?

250 × 1/1000 = 0.25dm3

Hence, g/dm3 = 6.43/0.25 = 25.72g

Mass Concentration = Molar Concentration × Molar mass

2 × 23 + 12 + 16 × 3 + 18x = 142.099

106 + 18x = 142.099

18x = 142.099 - 106

18x = 36.099

x = 36.099/18 = 2.0055

After obtaining the molar concentration to be 0.181moldm-3 and mass concentration to be 25.72gdm-3, we can find out the number of moles of water of crystallization following the steps below:

Hence, the number of moles of water of crystallization in the hydrated salt is 2. That's

**Example 2**: A student heated 4.38g of hydrated zinc tetraoxosulphate (VI) ZnSO4.xH2O and obtained 2.46g of anhydrous zinc tetraoxosulphate. Calculate the value of X. [Zn =65.4, S = 32, O = 16]

**Solution**

Mass of water of crystallization = 4.38 - 2.46

= 1.92g

Molar mass of ZnSO4 = 65.4 + 32 + 16 × 4 = 161.4g/mol

|  |  |  |
| --- | --- | --- |
|  | **ZnSO4** | **H2O** |
| **Mass** | **2.46** | **1.92** |
| **Molar mass** | **161.4** | **18** |
| **Mole** | **2.46/161.4 = 0.0152** | **1.92/18 = 0.10666** |
| **Divide through by the smallest mole ratio** | **0.0152/0.0152 = 1.0** | **0.10667/0.0152 = 7.017 ~ 7.0** |
|  | **1** | **7** |

Hence, the value of X in ZnSO4.xH2O is 7 and the formula becomes ZnSO4.7H2O. The name is zinc tetraoxosulphate (VI) heptahydrate.

**Example 3:** If 28.6g of hydrated sodium trioxocarbonate (IV) Na2CO3.XH2O was heated to 10.6g of anhydrous sodium trioxocarbonate. Calculate the value of X.

**Solution**

Na2CO3.XH2O ==> Na2CO3+ xH2O

28.6g 10.6g

Therefore, mass of steam (mass of water of crystallization) = Mass of hydrated – Mass of anhydrous = 28.6 - 10.6 = 18.0g H2O

Molar mass of Na2CO3 = 23 × 2 + 12 + 16 × 3 = 106g/mol

Molar mass of H2O = 1 × 2 + 16 = 18.0gmol-1

Find the mole ratio by dividing through by the smallest mole:

**Hence, the value of X in Na2CO3.xH2O is 10 and the formula becomes Na2CO3.10H2O. The name is sodium trioxocarbonate (IV) decahydrate.**

**Alternatively, using the formula;**

Mass of steam (mass of water of crystallization) =28.6 - 10.6 = 18.0gH2O

Cross multiplying, we get;

10.6 × 18x = 106 × 18

190.8x = 1908

Dividing through by 190.8, we have;

**Hence, the value of X in Na2CO3.xH2O is 10 and the formula becomes Na2CO3.10H2O. The name is sodium trioxocarbonate (IV) decahydrate.**

**Example 4:** 5.0g of a hydrated salt of barium when heated to a constant mass gave 4.26g of anhydrous salt with a molecular mass of 208gmol-1. Calculate the number of molecules of water of crystallization in the hydrated salt.

**Solution**

Mass of water of crystallization = 5.0 - 4.26 = 0.74g

Molar mass of water (H2O) = 1 × 2 + 16 = 18g/mol

Find the mole ratio by dividing through by the smallest mole:

Hence, number of molecules of water of crystallization in the hydrated salt is 2.

**Example 5:** A Sodium trioxocarbonate (IV) crystal, Na2CO3.xH2O contains 62.94% by mass of water of crystallization. Find the value of X.

**Solution**

The crystals contain 62.94% H2O and the rest (100 - 62.94 = 37.06%) must be anhydrous sodium trioxocarbonate (IV). That is 37.06 % of Na2CO3.

Molar mass of xH2O = X(1 × 2 + 16) = 18x g/mol

Molar mass of anhydrous Na2CO3 = 23 × 2 + 12 + 16 × 3 = 106 g/mol.

|  | Na2CO3 | H2O |
| --- | --- | --- |
| Mass | 37.06 | 62.94 |
| Molar mass | 106 | 18 |
| Mole | 37.06/106 = 0.349 | 62.94/18 = 3.49 |
| Divide through by the smallest mole ratio | 0.349/0.349 = 1.0 | 3.49/0.349 = 10 |
|  | 1 | 10 |

Hence, the formula for the compound is Na2CO3.10H2O and the name is sodium trioxocarbonate (IV) decahydrate.

**Alternatively, using the formula;**

Cross multiplying; 106 × 62.94 = 37.06 × 18x

6671.64 = 667.08x

Dividing through by 667.08, we get;

X = 10.00125 ~ 10

**Example 6:**

D is a solution of 0.100 moldm-3 HNO3. E is a solution containing 13.6 g of Na2CO3.yH2O per dm3.

Put D into the burette and titrate it against 20 cm3 or 25 cm3 portions of E using methyl orange as indicator. Repeat the titration to obtain consistent titres. Tabulate your burette readings and calculate the average volume of D used. The equation for the Reaction is:

2HNO3(aq) + Na2CO3.yH2O(aq) => Na2CO3(aq) + CO2(g) + H2O(l)

b) From your results and information provided above, calculate the:

i) Concentration of E in moldm-3;

ii) Concentration of E in gdm-3;

iii) Value of y in Na2CO3.yH2O.

[H= 1.00, C = 12.0, O = 16.0, Na = 23]

**Solution**

Volume of pipette = 25.0 cm3, Indicator used: Methyl orange

| Burette Reading | Rough (cm3) | 1st titre(cm3) | 2nd titre(cm3) |
| --- | --- | --- | --- |
| Final reading | 23.90 | 23.80 | 23.80 |
| Initial reading | 0.00 | 0.00 | 0.00 |
| Volume of acid used | 23.90 | 23.80 | 23.80 |

2HNO3(aq) + Na2CO3.yH2O(aq) => Na2CO3(aq) + CO2(g) + H2O(l)

bi) Concentration of E in moldm-3:

bii) Mass Concentration (gdm-3) = Molar Concentration(moldm-3) × Molar mass (g/mol)

Mass Concentration = 0.05 × 106 = 5.3gdm-3

biii) Value of y in Na2CO3.yH2O:

Mass of water of crystallization = 13.6 - 5.3

= 8.3g

**DILUTION FACTOR**

means to reduce the concentration of a solution.

is produced when a solute dissolves in a solvent.

refers to the amount of solute dissolved in a volume of solution.

**A solution can be diluted of by adding more solvent to the solution, then;**

1. Volume of solution increases from initial volume to final volume (Vintial to Vfinal). **The concentration decreases as the volume increases. C = k/V**
2. Concentration of solution decreases from initial concentration to final concentration (Cinitial to Cfinal).

refers to the ratio of the volume of the final (initial) solution to the volume of the initial (concentrated) solution, that is, the ratio of V2 to V1.

or

V2:V1

Supposed you are asked to prepare a 1:50 dilution of the solution. What it means is, take a known volume of the stock solution (Vinitial) and add enough solvent to it so that the solution has a new volume, Vfinal.

The"1:50"tells you the dilution factor, the ratio of volumes, to use to prepare the new solution.

V1 : V2

1 : 50

In this case, it tells us that V1 = 1 and V2 = 50, so the dilution factor, DF, =V2 ÷V1= 50 ÷ 1 = 50

That is, the new, diluted solution will have a volume 50 times greater than the volume of the original, undiluted, solution.

**You can use the equation:** Vfinal = DF × Vinitial to find the final volume of solution after dilution if you know the initial volume of the solution.

The dilution factor (DF) can be used alone or as the denominator of the fraction, for example, a DF of 10 means a 1:10 dilution, or 1 part solute + 9 parts diluents, for a total of 10 parts.

Assuming, you want to make 300 μL of a 1:250 dilution factor.

Substitute the values:

Rearrange:

**To make 300μL, measure out 1.2 μL of the stock solution and add water to make it 300 μL.**

**300μL–1.2 μL = 298.8 μL diluents volume (298.8μL of water will be added to 1.2 μL).**

**By rearranging the above equation, we get;**

**EXAMPLE 1:**

What is the dilution factor if you add a 0.1mL aliquot of a specimen to 9.9 mL of diluent?

**Solution**:

Vf = aliquot volume + diluents volume

= (0.1 + 9.9)mL = 10.0mL

DF = 10/0.1 =100

The dilution factor is often used as the denominator of a fraction.

For example, a DF of 100 means a 1: 100 dilution.

**EXAMPLE 2:**

How would you make 500 mL of a 1 : 250dilution?

**Solution:**

**Example 3:**

Given a stock solution of sodium chloride of 2.0mol/dm3, how would you prepare 250 cm3 of a 0.5mol/dm3 solution?

**Solution**

The required 0.5mol/dm3 concentration is ¼ of the original concentration of 2.0mol/dm3.

To make 1dm3 (1000cm3) of a 0.5mol/dm3 solution you would take 250cm3 of the stock solution and add 750cm3 of water.

Therefore to make only 250cm3 of solution you would mix ¼ of the above quantities i.e. mix 62.5cm3 of the stock solution plus187.5cm3 of pure water.

**Example 5:** You are given a stock solution of concentrated ammonia with a concentration of 17.9 moldm-3 (conc. ammonia! ~18M).

(a) What volume of the concentrated ammonia is needed to make up 1dm3 of 1.0 molar ammonia solution?

(b)What volume of concentrated ammonia is needed to make 5dm3 of a 1.5 molar solution?

**Solution**

1. Using dilution principle:

C1V1 = C2V2

17.90 × V1 = 1.0 × 1000

V1 = 1.0 × 1000/17.9 = 55.86

V1 = 55.9 cm3

b) C1V1 = C2V2

17.9 × V1 =1.5 × 5

V1 = 1.5 × 5/17.9

V1= 0.4189 ~ 0.419 dm3 or 419cm3

**Example 6:** A solution of HCl is standardized and found to be 1.183 M (moldm3). Calculate the volume of this solution that should be diluted to 1dm3 in a volumetric flask to produce a solution of HCl acid that is exactly 0.100 M.

**Solution**

Using dilution principle:

C1V1 = C2V2

1.183 × V1 = 0.1 × 1000

1.183 × V1 =100

V1 = 100/1.183

V1 = 84.53cm3 (0.08453 dm3)

**Example 7**

What volume of diluents must I add to 1mL of a 4mg/L solution of dexamethasone to get a concentration of 2.5 mg/L?

**Solution**

You can use the formula:

C1V1 = C2V2

C1 = 4mg/mL;V1 =1mL;

C2 = 2.5 mg/mL;V2 =?

V2 = V1 × C1/C2 = 1 × 4/2.5

= 1.6 mL

**You started with 1 mL and want 1.6 mL of diluted solution, so you must add (1.6 -1) mL = 0.6 mL of diluent. You will need 0.6 mL of diluent.**

***Then, to calculate the amount of water needed, use the following formula:***

***Final volume - Initial volume = volume of diluent***

**Example 8**: How would you prepare 500 ml of a 10% NaCl solution?

**Solution**

In this problem, the % solution is the number of grams solute in 100 ml solvent, so a 10% solution of NaCl is 10grams NaCl in 100 ml water. But you require 500ml, final volume.

10 % of 500g = 0.1 × 500 = 50 g NaCl

**TAKE AWAY**

1. In an experiment, 20.0cm3 portions of 0.065 mol dm-3NaOH were titrated against dilute HCl. The average volume of acid used is 23.45 cm3.

(a) (i) Name a suitable indicator for the titration. Give a reason for your answer.

(ii) Give the colour of the indicator in the base and at the end point.

(iii) What type of reaction is demonstrated by the experiment?

(b) (i) Write a balanced equation for the reaction.

(ii) Determine the average volume of acid used.

(c) Calculate the

(i) concentration of the acid in moldm-3 (ii) concentration of the acid in g dm-3

(iii) mass of HCl in 20cm3 of solution

[H = 1.00, Cl = 35.5 ]

**Answers:**

(a) (i) Methyl orange/methl red/phenolphthalein.

Because the end point will concise with the pH/colour change range of the indicator. It is a reaction between strong acid and a strong base.

(ii)

Colour in Base Colour at the End Point

\*Methyl orange Yellow Orange

\* Methyl red Yellow Orange

\* Phenolphtalein Pink Colourless

(iii) Neutralization

c)(i) CA = 0.0554moldm-3

(ii) Cone. in gdm -3 = 2.02 g dm-3

1. Mass of acid = 0.0404g
2. E was prepared by dissolving 33.60 g of a mixture of FeSO4.7H2O and NaCl to make 1.0dm3 solution. F was 0.0200 moldm-3 KMnO4. 25cm3of E was pipette into a conical flask acidified with 20.0cm3of dil. H2SO4 and titrated with F. Assuming, the average volume of F is 22.55 cm3.

From your results and the information provided, calculate the:

a)(i) concentration of FeSO4.7H2O in E in moldm-3;

(ii) Concentration of FeSO4.7H2O in E in gdm-3;

(iii) Percentage by mass of FeSO4.7H2O in the mixture.

[O = 16.0;S =32.0;H = 1.0;Fe = 56.0]

b)(i) Explain briefly why hydrochloric acid cannot be used to provide an acidic medium in titrations involving KMnO4.

(ii) Explain why the upper rim of the meniscus is read in KMnO4 titrations.

(iii)Why was no indicator used in the titration?

**Answers:**

**aI) 0.0902 moldm-3**

**II) 25.08 gdm-3**

**III) 74.7%**

**b(i) Hydrochloric acid cannot be used because KMnO4 will react with it to produce chlorine gas / KMnO4 will oxidise HCl(aq) to produce chlorine gas OR HCl (aq) will reduce KMnO4 to produce Mn2+**

**(ii) KMnO4 is a coloured solution hence, the lower meniscus cannot be seen while the upper meniscus can be seen and read.**

**(iii) No indicator was used because KMnO4 is a self-indicator / self – indicating/ acts or serves as an indicator.**

1. What is the Concentration of a solution which contains 0.28 g of potassium hydroxide in 100 cm3solution? [KOH = 56]. **Ans. = 0.05 moldm-3.**
2. 2.86 g of washing soda, Na2CO3.XH2O, were dissolved in 100 cm3 of water. 12.15 cm3 of this solution neutralized in 24.5 cm3 of 0.1 moldm-3, Calculate,

I) Molarity of the base,

II) % of water of crystallization,

III) The number of molecules of crystallization in the trioxocarbonate (IV).

**Ans. I) 0.098 moldm-3**

**II) 63%**

**III) 10 molecules**

1. What volume in cm3 of 0.2 moldm-3 HCl will neutralize 25 cm3 of 0.2 moldm-3 NaOH? **Ans. 25 cm3**

**CHAPTER 10: ELECTROLYSIS CALCULATIONS**

**INTRODUCTION**

**ELECTROLYSIS**

**Electro**  **lysis**

**FARADAY'S LAWS OF ELECTROLYSIS**

**Faraday'S Lirst Law of Electrolysis** states that st Law of Electrolysis𝑡𝑢𝑡𝑖𝑜𝑛𝑒𝑏𝑦 arbonate (IV).assium hydroxide in 100 cme seen and read.s a solution containing 13.6 g of Naoles of water.d. be more

m Q ----------(1)

Where, m = mass of a substance (in grams) deposited or liberated at electrode.

Q = amount of charge (in coulombs) or electricity passed through it.

On removing the proportionality in equation (1), we have;

M = ZQ

Where Z is the proportionality constant. Its unit is grams per coulomb (g/C). It is also called the electrochemical equivalent.

**Faraday is the proportionality consta**states that he proportionality constant. Its unit is grams per coulomb (g/C). It is also called the electrochemical equivalent. ining 13.6 g of Naoles of water.d. be “when the same quantity of electricity is passed through several electrolytes, the mass of the substances deposited are proportional to their respective chemical equivalent or equivalent weight”. Mathematically, it can be represented as follows –

w E

Where w = mass of the substance

E = equivalent weight of the substance.

**Equivalent weight or chemical equivalent of a substance can be defined as the ratio of its atomic weight and valency (number of electrons transferred).**

**e. g Cu2+ Cu;**

**Faraday𝑔 Second Law of Electrolysis can be further explained by the following example:**

Consider three different chemical reactions occurring in three separate electrolytic cells which are connected in series. Suppose in the 1st electrolytic cell sodium ion gains electrons and converts into sodium.

Na++ e−→ e

In 2nd electrolytic cell following reaction occurs:

Cu2++ 2e- → Cu

In 3rd electrolytic cell, the following reaction occurs l

Al3+ + 3e-  → Al

Suppose y moles of electrons are passed through three cells, the mass of sodium, aluminium and copper liberated are 23y grams, 9y grams, 31.75y grams respectively.

One mole of electrons is required for the reduction of one mole of ions. As we know, Charge on one electron is equal to 1.6021ns −19C and one mole of electron is equal to 6.023duc23 electrons. So, charge on one mole of electrons is equal to:

***(6.023 × 1023) × (1.6021 × 10−19C) = 96500 C***

**This charge (96500 C) is called 1 Faraday.**

If we pass 1 Faraday of charge in an electrolytic cell, then 1gm of equivalent weight of the substance will get deposited. So, we can write –

On combining the 1st and 2nd law we get:

Electric charge (Q) = current × time

1mole of electron = 1F

1F = 96500C

Q = I × t

Where **n** is number of moles of electrons, **m** is the reacting mass, **M** is the atomic mass of the element, **I** is the current, **t** is the time, **C** is the charge on the ion.

**Units:**

Q – coulombs (C), I - current (A) and t-time(S).

Where I = current, F = Faraday’s constant, m = reacting mass, t = time and M = atomic mass.

**Note: Always convert the time to seconds.**

**Solved Example 1:** A current of 2.0A flows through an electrolyte in 5 minutes, calculate the charge transferred during the electrolysis.

**Solution:**

Q = I × t

Convert time to seconds

60seconds = 1 minutes

? = 5minutes

Therefore 5 × 60 = 300

Q = I × t

= 2 × 300

= 600C

**Solved Example 2:** Some Lead (ii) bromide is heated in a crucible until it is molten. Two carbon electrodes are inserted in the crucible and an electric current of 1.8A flows through the molten lead (ii) bromide for 12 minutes. Calculate the mass of lead deposited at the cathode.(Pb = 207)

**Solution:**

Q = I × t

1minute = 60sec

12 min =?

= 1.8 × 12 × 60

= 1,296C

The next step is to write the equation for the electrolysis of lead (ii) bromide

Pb2++2e-Pb(s)

2F are required to produce 1 mole of lead (from the equation).

2 × 96500 of charge is required to produce 207g of lead.

1,296 C of charge would be required to produce;

2F = 207g of Pb

1296 C = ?

**Example 3:** Dilute sulphuric acid is electrolyzed using graphite electrodes. A current of 0.8A flows through the solution for a period of 30 minutes. Calculate the volume of oxygen gas produced at room temperature and pressure. (Faraday’s constant = 96500C, 1 mole of a gas occupies 22.4dm3 at s.t.p)

**Solution:**

Q = I × t

= 0.8 × 30 × 60 = 1440 C

The next step is to write the half-equation for the reaction:

## This is a very important equation that you have to commend to memory.

**4F is required to produce 1 mole of oxygen gas.**

4 × 96500C produces 22.4dm3 (r.t .p)

1440C produces;

Recall: 1 mole = 22.4dm3 at room temperature

0.08356 mole = ?

**Example 4:** A metal spoon is being copper plated using a copper (II) sulphate solution. A current of 0.5Aflows through the solution for a period of 2 hours. Calculate the mass of copper deposited on the spoon. (1F = 96500, relative atomic mass of copper = 64).

**Solution:**

Q = It

= 0.5 × 2 × 60 × 60

= 3600

The next step is to write a half equation;

2F are required to produce 1 mole of copper

(2 × 96500) produces 64g of copper

36000C will produce;

That is;

2 × 96500C = 64g Cu

36000C = ? of Cu

Cross multiplying we get;

**Note: In electrolysis calculations, it is important that you commend to memory the different half reactions. This will help you to know the number of moles of electrons required. For instance, Oxygen is often the anode product from the electrolysis of an aqueous solution. There are two common ways of writing an equation for its formation. One way is to show hydroxide ions losing electrons to form water:**

4OH-→ 2H2O + O2 + 4e-

**The other way is water gaining electrons;**

2H2O + 2e-------->H2 + 2OH-

**For chlorine, 2 moles of electrons are involved.**

2Cl-(aq) --------->Cl2(g) + 2e-

**Example 5:** A battery delivers 0.300A for 15min what amount of electrons in moles is transferred?

**Solution:**

Q = It

n =

Q = nF

Q = = 270 C

= 2.8×10-3

**Example 6:** What current in ampere will deposit 0.27g of aluminium in 2 hours? [Al = 27, F = 96500 C]

**Solution:**

**MARATHON SOLVED QUESTIONS**

**Example 1.** Electrolysis of dilute aqueous NaCl solution was carried out by passing 10 milliampere current. What is the time required to liberate 0.01mol of H2 gas at the cathode?

**Solution**

According to Faraday's first law of electrolysis:

m = EIt/F

Where:

M = mass of substance

E = equivalent weight of substance

I = current in amperes

T = time required

F = 96,500Coulombs.

2H+ + 2e−→H2

Mass of H2 = m = 0.01mol x 2gmol-1 = 0.02g

Equivalent weight of H2= E = Atomic weight/valence =1gmol-1/1 =1gmol-1

Current in ampere = I = 10milliamperes = 10 x 10-3 ampere.

t = mF/EI

t = (0.01 × 2) × 96500/1 × 10 × 10-3

t = 19.3

**Example 2.** A 4.0 molar aqueous solution of NaCl is prepared and 500 mL of this solution is electrolyzed. This leads to the evolution of chlorine gas at one of electrodes (relative atomic mass of Na = 23, Hg = 200; 1 Faraday constant = 96500 Coulombs mol-1):

i) The total number of moles of chlorine gas evolved is:

ii) If the cathode is Hg electrode, the maximum mass of Amalgam formed from this solution is:

**Solution:**

\* First, we have to calculate the amount of NaCl present in 4.0 molar solutions.

The number of moles of NaCl in 500 mL of 4.0 molar solution = Molarity x Volume (in L) = 4.0 x 500 x 10-3 = 2.0 moles.

\* The following reaction occurs during electrolysis of aqueous solution of NaCl.

Dissociation of NaCl: 2NaCl(aq) ------> 2Na+(aq) + 2Cl-(aq)

At cathode: 2H2O + 2e- -------> H2 + 2OH-

At anode: 2Cl-(aq) ---------> Cl2(g) + 2e-

Complete reaction: 2NaCl(aq) + 2H2O ------>H2 + Cl2(g) + 2Na+(aq) + 2OH-(aq)

That means, two moles of NaCl gives one mole of Cl2 gas.

Since there are 2.0 moles of NaCl present in the solution, one mole of Cl2 gas will be evolved at anode upon complete electrolysis.

ii) If the cathode is Hg electrode, the maximum mass of Amalgam formed from this solution is:

\* When Hg is used as cathode, Na+ will be reduced to Na instead of H2O. Thus formed Na will form sodium amalgam, Na-Hg.

At cathode: Na+ + e- + Hg -------> Na-Hg

2.0 moles of NaCl present in the solution contain 2 moles of Na+. Hence two moles of Na-Hg is formed.

The mass of 2 moles of Na-Hg = 2 (23 + 200) = 446g.

3) The amount of electricity that can deposit 108 g of silver from AgNO3 solution is

**Solution**

For silver, the gram equivalent weight is 108 g. We know that 1 Faraday of electricity is required for deposition of 1 gram equivalent weight of an element.

Ag+ + e- ==> Ag

One mole of electron equals one Faraday. Hence, amount of electricity deposited is 1F.

4) During the electrolysis of molten sodium chloride, the time required to produce 0.10 mol of chlorine gas using a current of 3 amperes is

**Solution:**

Mass of chlorine gas = m = no. of moles x molecular weight = 0.10 mol x 71 g

Equivalent weight of Cl2 gas = E = 35.5/1 = 35.5 g

t = mF/EI

t = (0.10 35.5/1 = 35.5 g= no. o

t = 6433 seconds

60 seconds = 1 minute

6433 seconds = ?

Therefore, time required for deposition of chlorine gas is 107.22 mins.

5) During the electrolysis of molten Al2O3, a current of 6A was passed through the electrolyte for 1 hr 30mins. Calculate the mass of aluminium deposited at the cathode.

**Solution**

Q = I x t

= 6 x (30 + 60) x 60

= 32400 C

Al3+ + 3e- → Al(s)

3F of aluminium = 27g

32400C of aluminium = ?

27g 0C of aluminium = ? cat

**Recall: 1F = 96500. Therefore, 3F = 3 × 96500 = 289500C.**

6) Calculate the time it will take to deposit 1 mole of silver, if a current of 6A is passed through a solution of silver trioxonitrate (V). [IF = 96500C]

**Solution**

Ag+(aq) + e- ==> Ag (s)

1 mole of electrons = 1F = 96500C

96500C deposit 1 mole of Ag

Q = It

t = 96500C /6A

= 16083.3 seconds

7. In the electrolysis of copper sulphate solution using carbon electrodes, how many moles of Cu is deposited, what mass and volume of oxygen would be formed at the positive electrode if 254g of copper was deposited on the negative electrode? [Atomic masses: Cu = 63.5, O = 16].

Solution

Cu2+(aq)tiona- aq)tio(s) s))tion-(aq) aq)t-a==>)ti2O(l))+ O2(g)

It takes a transfer of 2 moles of electrons to form 1 mole of solid copper (63.5g) from 1 mole of copper(II) ions, Cu2+and a transfer of 4 moles of electrons to form 1 mole of oxygen from 4 moles of hydroxide, OH- nd a

Therefore the expected mole ratio of Cu(s) s)r2(g) (g)efore the expected mole rat

So moles of oxygen formed = 1 mole, the atomic mass of O2 = 32g. Hence, the mass of oxygen formed = 32g and volume of 1 mole of oxygen =12 x 22.4 = 22.4 dm3d

8. In the industrial manufacture of aluminium by electrolysis of the molten oxide (plus cryolite) 250kg of aluminium are formed. What volume of oxygen would be theoretically formed at room temperature and pressure?[ Ar(Al) = 27 and 1 mole of gas at R.T.P = 22.4 dm3 (litres) ]

**Solution**

Aluminium oxide is Al2O3, so on splitting in electrolysis, the atomic ratio forelectrolysiseoretically formed at room temp2torhe ato

4Al3+(aq) aq) a-a==>) at(s) s)) at2-(aq) - 12e- ==>2ea2(g)

**Note:** It takes 12 electrons added to four Al3+ +t takes 12 electrons added to four Alcally formed at room temperoxide ions, O2-, to form six oxygen atoms, which combine to form three O2 molecules.

BUT, oxygen exists as O2Umolecules, so the mole ratio ofcAl atoms: O2 molecules, so the .

Convert kg to g;

250kg Al = 250000g,

From the equation, 4 moles of Al = 3 moles of O2

Therefore, 9259.26 moles of Al = ?

Since volume of 1 mole of gas at RTP = 22.4 dm3 dm of 1

6944.445 moles of O2 =m?

9. A current was passed through an electrolysis circuit of silver nitrate solution and O.54g of silver was formed. Ar(Ag) = 108 and the electrode equation is Ag+  108 and- ==>08 (s). Ar(Ag) = 64 and the electrode equation iss A2++64 a+ 2e- ==>e a(s). If in the same circuit a copper (II) sulphate and copper electrodes cell was connected, how much copper is deposited at the negative (-) cathode?

**Solution**

0.54g Ag = 0.54 / 108 = 0.005 mol Ag

1 mole electron deposits 1 mole of silver.

In the same circuit;

If 2 moles of electrons of Cu = 0.005 mole of Cu

1 mole of electron of Cu = ?

10. How much copper is deposited if a current of 0.2 Amps is passed for 2 hours through a copper (II) sulphate solution?

**Solution**

Electrode equation: (-) cathode: le2+(aq)t+ 2e- ==>etr(s) s)etrode equati

The quantity of electricity passed in Coulombs = current in A x time in seconds (Q = I x t)

= 0.2 x 2 x 60 x 60 = 1440 Coulombs, and 1 mole of electron = 96500 Coulombs.

2 moles of electrons = 1 mole of copper deposited

0.01492 moles of electrons = ?

Mass of copper = moles of copper x atomic mass of copper

Mass of copper = 0.00746 x 64 =r x atomic mass of copperted.

11. In the electrolysis of molten sodium chloride 60 cm3 1. In the electrolysis o

Electrode equations:

(-) cathodequa+ -) - -) cathodequa+ -) c- -) cathodequations:ysis of molten sodiu

(+) cathodequ-+-2e-2==> ca2=

Calculate:

(a) how many moles of were chlorine produced?

(b) what mass of sodium would be formed?

(c) for how long would a current of 3 A in the electrolysis circuit have to flow to produce the 60cm3 c) for how l

**Solution**

a) Mole = volume/molar volume = 60/24000 = 0.0025 mol of Cl2

b) From the electrode equations 2mol sodium will be made for every mole of chlorine

So, 0.0025 x 2 = 0.005mol sodium will be formed. Ar (Na) = 23

Mass = atomic mass or formula mass = 0.005 x 23 = 0.115gNa

c) To produce 0.0025 mole of Cl2 you need 0.005 moles of electrons

0.005 mole electrons = 0.005 x 96500 coulombs = 482.5C

Q = I x t, so 482.5 = 2x t, therefore t = 482.5/3 = 161s (to nearest second)

**TAKE AWAY**

1. Calculate the quantity of electricity obtained from 2 moles of electrons. **Ans. 193,000C**

2. Calculate the moles of electrons obtained from 250 C of electricity. **Ans.2.59 mol**

3. How many coulombs of electricity would liberate 1.08 g of Ag from a solution of silver salt? [Ag = 108.0; 1F = 96500 C].  **Ans. = 965 C**

4. What is the mass of silver in grams deposited when a current of 2A is passed through a solution of silver salt for 10 minutes?

[Ag = 108, 1F = 96500C] **Ans. 1.34g**

CHAPTER 11: OXIDATION NUMBERS/STATES

OXIDATION NUMBERS/STATES

INTRODUCTION

In general, oxidation state or number helps us describe the transfer of electrons. However, students have to note that it is different from a formal charge which determines the arrangement of atoms. The oxidation number or state is also used to determine the changes that occur in redox reactions. Mean while, it is quite similar to valence electrons.

**What is Oxidation Number?**

Oxidation number of an atom is defined as the charge that an atom appears to have on forming ionic bonds with other heteroatom. An atom having higher electronegativity (even if it forms a covalent bond) is given a negative oxidation state.

**Rules for Working out oxidation states**

You do not work out oxidation states by counting the numbers of electrons transferred. It would take far too long. Instead you learn some simple rules, and do some very simple sums.

1. The oxidation state of an uncombined element is zero. That is obviously so, because it has not been either oxidized or reduced yet. This applies whatever the structure of the element-whether it is, for example, Xe or Cl2 or S8, or whether it has a giant structure like carbon or silicon.

2. The sum of the oxidation states of all the atoms or ions in a neutral compound is zero.

The sum of the oxidation states of all the atoms in anion is equal to the charge on the ion.

3. The more electronegative element in a substance is given a negative oxidation state. The less electronegative one is given a positive oxidation state. Remember that fluorine is the most electronegative element with oxygen second.

4. Some elements almost always have the same oxidation states in their compounds:

* Group 1 metals always have +1.
* Group 2 metals always have +2.
* Oxygen usually has -2 except in peroxides and F2O.
* Hydrogen usually have +1 except in metal hydrides where it -1.
* Fluorine always has -1.
* Chlorine usually has -1 except in compounds with O or F.

**THE REASONS FOR THE EXCEPTIONS**

**Hydrogen in the metal hydrides**

Metal hydrides include compounds like sodium hydride, NaH. In this, the hydrogen is present as a hydride ion, H-. The oxidation state of a simple ion like hydride is equal to the charge on the ion. In this case, -1.

Alternatively, you can think of it that the sum of the oxidation states in a neutral compound is zero. Since Group 1 metals always have an oxidation state of +1 in their compounds, it follows that the hydrogen must have an oxidation state of -1 (+ 1 – 1 = 0).

**Oxygen in peroxides**

Peroxides include hydrogen peroxide, H2O2. This is an electrically neutral compound and so the sum of the oxidation states of the hydrogen and oxygen must be zero.

Since each hydrogen has an oxidation state of +1, each oxygen must have an oxidation state of -1 to balance it.

**Oxygen in F2O**

The problem here is that oxygen is not the most electronegative element. The fluorine is more electronegative and has an oxidation state of -1. In this case, the oxygen has an oxidation state of +2.

**Chlorine in compounds with fluorine or oxygen**

There are so many different oxidation states that chlorine can have in these, that it is safer to simply remember that the chlorine does not have an oxidation state of -1 in them, and work out its actual oxidation state when you need it.

**CALCULATING THE OXIDATION NUMBERS/STATES**

**Example 1:** What is the oxidation state of chromium in Cr2+?

**Solution**

That is easy. For a simple ion like this, the oxidation state is the charge on the ion-in other words: +2

**Example 2:** What is the oxidation state of chromium in CrCl3?

**Solution**

This is a neutral compound, so the sum of the oxidation states is zero. Chlorine has an oxidation state of -1. If the oxidation state of chromium is n:

n + 3( -1 ) = 0

n = +3

**Example 3:** Calculate the oxidation number of iron in [Fe(H2O)6]3+

**Solution**

The sum of the oxidation numbers of all constituent atoms must be equal to +3 (charge on top of the compound). Also, the oxidation number of H2Ois 0 since it is neutral ligand. If the oxidation number of iron is x then;x + (6 × 0) = +3

∴ x = +3

**Example 4:** Determine the oxidation number of platinum in [Pt(NH3)2Cl4].

**Solution**

Let the oxidation number of platinum be x. Since NH3 is a neutral ligand, then;

x + (2 × 0) + (4 × {–1}) = 0

So, x – 4 = 0

∴ x = +4

**Example 5:** What is the oxidation state of chromium in Cr(H2O)63+?

**Solution**

This is an ion and so the sum of the oxidation states is equal to the charge on the ion. There is a short-cut for working out oxidation states in complex ions, like this where the metal atom is surrounded by electrically neutral molecules like water or ammonia (ligand). The sum of the oxidation states in the attached neutral molecule must be zero. That means that you can ignore them when you do the sum. This would be essentially the same as an unattached chromium ion, Cr3+. The oxidation state is +3.

**Example 6:** What is the oxidation state of chromium in the dichromate ion, Cr2O72-?

**Solution**

The oxidation state of the oxygen is -2, and the sum of the oxidation states is equal to the charge on the ion. Do not forget that there are 2 chromium atoms present.

2n + 7(-2) = -2

2n – 14 = -2

2n = -2 + 14

2n = 12

Dividing both sides by 2, we get;

n = +6

**Examples 7**: Oxidation state of chlorine in KCl. KCl is neutral and so, net charge = 0

**Solution**

Oxidation state of KCl = Oxidation state of potassium + oxidation state of chlorine = 0.

Let the oxidation state of chlorine = x

Oxidation state of potassium = +1

Sum of Oxidation states: +1 + x = 0: x = -1

Oxidation state of chlorine in KCl = -1

**Example 8**: Oxidation number of Manganese in permanganate ion MnO4–

**Solution**

Charge on the permanganate ion is -1

Oxidation state of permanganate ion = Oxidation state of manganese + 4 oxidation state of oxygen = -1.

Oxidation state of oxygen = -2

Sum of Oxidation states: x + (4 × -2) = -1:

X – 8 = -1

X = -1 + 8 = +7

Oxidation state of manganese = +7

**Example 9:** Oxidation number of a metal ion in a complex.

1. Ni(CO)4.
2. [CoCl2(NH3)4]Cl.

**Solution**

1. The total charge of the complex is zero. CO is a neutral molecule.

Oxidation states: x +(4 × 0) = 0:x = 0

Nickel is also in zero oxidation state.

1. [CoCl2(NH3)4]+

Ammonia is a neutral ligand and chlorine has a unit negative charge.

Oxidation number of [CoCl2(NH3)4]+= Oxidation number of (Co +2Cl + 4 × 0) = +1.

Oxidation states → x +(2 × -1) + 4 × 0 = +1:

X - 2 + 0 = +1

X = +1 + 2 = +3

Oxidation number of cobalt in the complex = +3

**Example 10:** Calculate the oxidation numbers of chlorine in Cl2O, Cl2O5 and Cl2O7.

**Solution**

1. Cl2O:

Cl2O is neutral and so, net charge = 0.

Let the oxidation state of chlorine = x

Net oxidation state of Cl2O = 2x (Oxidation state of chlorine) +1 ×-2 (Oxidation state of oxygen = -2) = 0

Oxidation state of oxygen = -2.

Oxidation states: 2x +(1 × -2) = 0

2x – 2 = 0

2x = 2

​Dividing both sides by 2, we get;

X = +1

Hence, oxidation state of chlorine in Cl2O =+1

1. Cl2O5:

Cl2O5 is neutral and so, net charge = 0

Let the oxidation state of chlorine = x

Oxidation state of oxygen = -2

Oxidation state of Cl2O 5= 2x (Oxidation state of chlorine) + 5x-2(oxidation state of oxygen) = 0.

Oxidation states: 2x + (5× -2) = 0

2x – 10 = 0

2x = 10

​Dividing both sides by 2, we get;

X = +5

Oxidation state of chlorine in Cl2O5 = +5

1. Cl2O7:

Cl2O7 is neutral and so, net charge = 0

Let the oxidation state of chlorine = x

Oxidation state of oxygen = -2

Oxidation state of Cl2O7 = 2x (Oxidation state of chlorine) +7 x oxidation state of oxygen = 0.

Oxidation states: 2x +(7× -2) = 0

2x – 14 = 0

2x = 14

​Dividing both sides by 2, we get;

X = +7

Oxidation state of chlorine in Cl2O7 = +7

**CALCULATING CELL EMF (Electromotive force)**

A cell emf is a measure of the driving force of the cell reaction. This reaction occurs in the cell as separate half-reactions (an oxidation half-reaction and a reduction half-reaction. Mathematically, it is given as:

ECell = Ecathode - Eanode

* **Recall:**
* **That oxidation takes place at the anode while reduction takes place at the cathode.**

**E.g Zn/Zn2+ // Cu2+ /Cu**

**Zn => Zn2+ + 2e-(oxidation- loss of electrons)**

**Cu2+ + 2e- => Cu (reduction- gain of electrons)**

* **That oxidation half cell generate electrons which flow to the reduction half cell.**
* **That ions are being reduced while the metal (electrode) is being oxidized, so that an equilibrium is set up.Cu2+(aq)  + 2e- <==> Cu(s)**

**Solved Example 1:**

Suppose a voltaic cell is constructed from a cadmium electrode and a silver electrode having the following reduction half reactions and corresponding standard electrode potentials:

Cd2+(aq) + 2e- =>Cd(s): E Cd: 0.40V

Ag+(aq) + e- => Ag(s): E Ag: 0.80V

**SOLUTION:**

You have to reverse one of the half reactions to obtain the oxidation part of the cell reaction. The cell reaction is spontaneous with stronger reducing agent on the left (oxidation). This will be Cd since it is the reactant in the half reaction with more negative electrode potential. Hence, you reverse the first half reaction, as well as the sign of the half-cell potential.

Cd(s) => Cd2+(aq) + 2e-. : -Ehe sign of

Ag+(aq) + e- => Ag(s) : E 0.80V

To obtain the cell reaction, you have to balance the electron gain and loss. This is done by multiplying the half reactions by factors so that when the half reactions are added together, the electrons cancel. This does not affect the half-cell potentials.

We multiply the oxidation side by one electron gained in the reduction side and also multiply the reduction side by two electrons loss in the oxidation side.

1 × {Cd(s) => Cd2+(aq) + 2e- : -E

2 -E +(aq) +e- => Ag(s): E- => Ag

We move on to adding the two half - cells together:

Cd(s) =>Cd2+(aq) + 2e‑ : -E

2Ag+(aq) + 2e- =>2Ag(s) : E

Cd(s) + 2Ag+(aq) =>Cd2+(aq) + 2Ag(s)

E 1.20V {E

The cell notation is:

Cd(s) / Cd2+(aq) // Ag+(aq) / Ag(s)

**Note:**

**Example 2:**

Calculate the standard EMF of the following voltaic cell at 25𝒆 𝒆𝒍𝒆𝒄𝒕𝒓𝒐𝒅𝒆 𝒑𝒐𝒕𝒆𝒏𝒕𝒊𝒂.

Al3+(aq) + 3e- => Al(s): Ee standard

Fe2+(aq) + 2e- => Fe(s): Es))Al -0.41V

**Solution:**

You have to reverse the first half reaction and standard electrode potential since it is the strongest reducing agent (less value of electrode potential).

Al(s) =>Al3+(aq) + 3e- :E)have to re

Fe2+(aq) +2e- => Fe(s): E) Fev -0.41V

You now balance the electrical charges (electron loss and gain) by multiplying the first half-reaction by 2 and the second half reaction by 3, such that when the half reactions are added together, the electrons cancel.

2Al(s) =>3Al3+(aq) + 6e- :E) multiplyi

3Fe2+(aq) + 6e- =>3Fe(s): Eg the first

The addition of the two half-reactions gives:

2Al3+(s) + 3Fe2+(aq) =>2Al3+(aq) + 3Fe(s)

You can calculate the cell from the formula,

Eocell = Elcathode - Eoanode

Encell = El e calcu

= -0.41 -(-1.66)

= 1.26V

**The Reaction is spontaneous since the EMF is positive.**

The table below shows the standard electrode (reduction) potential in aqueous solution at 25sit

**Example 3:** Using standard electrode potentials, Calculate Eitive.st 25°C for the following cells:

1. Zn(s)/ Zn2+(aq)//Cu2+(aq)/Cu:
2. Zn(s)/Zn2+(aq)//Ag(s)/Ag+(aq)

Zn2+(aq) + 2e- => Zn(s) E -0.76V

Cu2+(aq) + 2e- => Cu(s) E Cu n

Ag+(aq) + e- => Ag (s) E = 0.80V

**SOLUTION:**

1. Zn(s) => Zn2+(aq) +2e- -EenI = 0.76

Cu2+(aq) + 2e- => Cu(s) E = 0.34V

The electrical charges are already balanced. Use the formula;

Ehcell = Elcathode -Etanode

Enodede t = 0.34V. Echarge = 0.76V

E 0.76V. Echarges are alr

= 0.34V - ( -0.76)

= 0.34 + 0.76

E.cell =1.10V. Hence spontaneous

2. Zn(s) => Zn2+(aq) +2e- -EenZ = 0.76V

Ag+(aq) + e- => Ag(s) E = 0.80V

**Balance the electrical charges (electron loss and gain) by multiplying the first half-reaction by one (1 electron gained by Ag+) and the second half reactions by two (2 electrons loss by Zn), such that when the half-reactions are added, the electrons cancel out.**

1[ Zn(s)=> Zn2+(aq) +2e- ]:-E Zn = 0.76V

2[Ag+(aq) + 2e- =>2Ag (s)]: EsAg+ = 0.80V

Zn(s)=> Zn2+(aq) +2e- :-EeZn = 0.76V

2Ag+(aq) + 2e- =>2Ag (s): EsAg+ = 0.80V

The addition of the two half-reactions gives:

Zn(s) + 2Ag+(aq)  => Zn2+ + 2Ag(s)

You can now calculate the EMF using the formula

Elculate thcathode - E-anode

= Edcathode -( -Edanode)

= 0.80V - ( -0.76)

= 0.80V + 0.76V

= 1.56V. Hence, spontaneous.

**DIFFERENCES BETWEEN VOLTAIC CELL AND ELECTROLYTIC CELL**

|  |  |
| --- | --- |
| **Voltaic cell** | **Electrolytic cell** |
| It is an electric cell | Require a battery |
| Anode is negative | Anode is positive |
| Cathode is positive | Cathode is negative |
| Produces electricity | Uses electricity |
| Spontaneous reaction | Non-spontaneous reaction |
| Two different elements (two different solutions | The same element(One Solution both electrodes) |

**TAKE AWAY**

1. Calculate the standard EMF of the following cell at 25°C:

Cr(s) / Cr3+(aq) // Hg22+(aq) /Hg(l)

**Ans: 1.54V**

1. Calculate the standard EMF of the following cell at 25°C:

Sn(s) / Sn2+(aq)// Cu2+(aq) /Cu(s)

**Ans: 0.30V**

1. What is the oxidation state of chlorine in Cl2O4? **Ans. +4**
2. Calculate the oxidation number of carbon in HCO3-  **Ans. +4**
3. Find the oxidation number of chlorine in NaClO3. **Ans. +5**

CHAPTER 12: ENERGY CHANGES

INTRODUCTION

* Enthalpy of Reaction (Heat of Reaction) is the heat liberated or the heat absorbed when a chemical reaction takes place.
* An exothermic reaction liberates heat, temperature of the reaction mixture increases.
* An endothermic reaction absorbs heat, temperature of the reaction mixture decreases.
* The units of enthalpy of reaction, or heat of reaction, are kJmol-1 for a specified reactant or product.
* The enthalpy of reaction (heat of reaction) for a neutralization reaction is known as the enthalpy of neutralization (heat of neutralization). For instance, the heat accompanying a reaction between an acid and a base to form salt and water.
* The enthalpy of reaction (heat of reaction) for a solute dissolving in a solvent is known as the enthalpy of solution (heat of solution).
* The enthalpy of reaction (heat of reaction) for a precipitation reaction is known as the enthalpy of precipitation (heat of precipitation).

**Enthalpy of reaction (heat of reaction) can be measured experimentally using a calorimeter.**

Typically, the calculation for heat released or absorbed, q, for the reaction of aqueous solutions is measured in unit of joules (J).

**Example 1:**

In an experiment, 1.2g of sodium hydroxide pellets, NaOH(s), were dissolved in 100mL of water at 25°C. The temperature of the water rose to 27.5°C. Calculate the enthalpy change (heat of solution) for the reaction in kJmol-1 of solute. Calculate the heat released, q, in joules (J), by the reaction.

**Solution**

Where q = heat loss or gained in joules. This the same as the enthalpy change if the pressure is constant m = mass of the substance, C = specific heat capacity and ∆T = change in temperature (Tf - Ti)

q = 100 × 4.184 × (27.5 - 25) = 1046 J

Calculate the moles of solute (NaOH):

Moles = mass ÷ molar mass

Moles (NaOH) = 1.2 ÷ (22.99 + 16.00 + 1.008) = 0.030 mol.

n(NaOH) = 0.030mol. Calculate the enthalpy change, ΔH, in kJmol-1 of solute:

ΔH = -q ÷ n (solute) but if you want it in KJ per mol, you have to divide it by 1000.

ΔH = -q/1000 ÷ n (solute) = -1046/1000 ÷ 0.030 = -35kJmol-1

ΔH is negative because the reaction is exothermic (energy is released causing the temperature of the solution to increase).

**Example 2:** When 3.02 g of NH4Cl is dissolved in enough water to make 20.05 mL of solution, the temperature dropped from 19.8°C to 9.1°C. Calculate the enthalpy change (in kJ) when 1 mole of NH4Cl is dissolved in water. The density and specific heat of water are 1.00 g/mL and 4.18 J/g°C respectively.

**Solution**

q = mC∆T

q = heat, m = mass = 20.05 ml x 1g/ml = 20.05 g + 3.02 g = 23.07 g (note: in some introductory courses, they use only the mass of water, i.e. 20.05 g, but in more advanced courses they will add the mass of solute), C = sp. heat = 4.18 J/g/deg, ∆T = change in temperature = 19.8 - 9.1 = 10.7 degree Celcius (Note: this is an endothermic reaction)

Solving for q:

q = (23.07 g)(4.18 J/g/deg)(10.7 deg) = 1032 J

This is the heat for dissolution of 3.02 g NH4Cl. The question asks for enthalpy change for 1 mole of NH4Cl.

If 0.056 moles = 1032J

1 mole of NH4Cl =?

**Enthalpy of Neutralization**

**Calorimetry**

The device used experimentally to determine the heat associated with a chemical reaction is called a calorimeter. Calorimetry, the science of measuring heat, is based on observing the temperature change when a body absorbs or discharges energy as heat. Substances respond differently to being heated. One substance might require a great deal of heat energy to raise its temperature by one degree, whereas another will exhibit the same temperature change after absorbing relatively little heat.

When an element or a compound is heated, the energy required will depend on the amount of the substance present (for example, it takes twice as much energy to raise the temperature of 2 g of water by one degree than it takes to raise the temperature of 1 g of water by one degree). Thus, in defining the heat capacity of a substance, the amount of substance must be specified. If the heat capacity is given per gram of substance, it is called the specific heat capacity, and its units are J/°C . g or J/K . g. If the heat capacity is given per mole of the substance, it is called the molar heat capacity, and it has the units J/°C . mol or J/K . mol. It takes much less energy to change the temperature of a gram of a metal by 18°C than for a gram of water.

Recall that under these conditions, the change in enthalpy equals the heat; that is, ∆H = qP.

For example, suppose we mix 50.0 mL of 1.0 M HCl at 25.0°C with 50.0 mL of 1.0 M NaOH also at 25°C in a calorimeter. After the reactants are mixed by stirring, the temperature is observed to increase to 31.9°C. The net ionic equation for this reaction is

H+(aq) +OH-(aq)=> H2O

When these reactants (each originally at the same temperature) are mixed, the temperature of the mixed solution is observed to increase. Therefore, the chemical reaction must be releasing energy as heat. This released energy increases the random motions of the solution components, which in turn increases the temperature. The quantity of energy released can be determined from the temperature increase, the mass of solution, and the specific heat capacity of the solution. For an approximate result, we will assume that the calorimeter does not absorb or leak any heat and that the solution can be treated as if it were pure water with a density of 1.0 g/mL. We also need to know the heat required to raise the temperature of a given amount of water by 1°C.

**What is meant by the specific heat capacity of water is 4.18 J/°C . g? This means that 4.18 J of energy is required to raise the temperature of 1 g of water by 1°C.**

**Example 2:**

In an experiment to determine the enthalpy change, ΔH, for the neutralization reaction:

NaOH(aq) + HCl(aq) → NaCl(aq) + H2O(l)

The following results were obtained:

Mass of 100 mL of 0.50 molL-1 HCl = Ma = 100g

Mass of 100 mL of 0.50 molL-1 NaOH = Mb = 100g

Initial Temperature = Ti = 20.1°C

Final Temperature = Tf =23.4°C

Specific heat capacity of solutions = C = 4.184J°C-1g-1

Calculate the:

1. Calculate the heat released, q, in Joules (J) by the neutralization reaction.
2. Enthalpy change, ΔH, in kJmol-1 of water formed for the reaction.

**Solution**

q = mc∆T

q = mass (reaction mixture) × specific heat capacity (water) × change in temperature (solution)

q = (Ma + Mb) × C × (Tf - Ti)

Ma and Mb are the masses of acid and base respectively.

q = (100 + 100) × 4.184 × (23.4 - 20.1)

= 200 × 4.184 × 3.3

= 2761.44 J. Since energy is released; the value will be negative i.e -2761.44J.

Calculate the moles of reactants: Moles = Molarity × Volume

Moles (NaOH) = 0.50molL-1 × (100 × 10-3)L = 0.05mol

Moles (HCl) = 0.50molL-1 × (100 × 10-3)L = 0.05mol

NaOH(aq) and HCl(aq) are in 1:1 mole ratio which is an exact stoichiometric ratio as shown by the neutralization equation.

0.05 mole of NaOH (aq) reacts with 0.05 mole HCl (aq) to produce 0.05 mole of water.

Calculate the enthalpy (heat) of reaction, ΔH, in kJmol-1;

Since the reactants are present in a 1:1 stoichiometric ratio, 0.05 mole of NaOH(aq) reacts with 0.05 mole HCl(aq) to produce 0.05 mole of water,

Moles (H2O(l)) = n(H2O(l)) = 0.05mol

or

ΔH = - q/1000 ÷ n(H2O(l))

ΔH is negative because the reaction is exothermic.

**Enthalpy of Precipitation (Heat of Precipitation)**

**Example 4:**

50mL of 0.20 molL-1 lead(II) nitrate solution, Pb(NO3)2(aq), at 19.6°C was added to 30mL of a solution containing excess potassium iodide, KI(aq) also at 19.6°C. The solutions reacted to form a yellow lead (II) iodide precipitate, PbI2(s), and the temperature of the reaction mixture increased to 22.2°C.

Pb(NO3)2(aq) + 2KI(aq) → PbI2(s) + 2KNO3(aq)

1. Calculate the heat released, q, in Joules (J), by the precipitation reaction.
2. Calculate the enthalpy change per mole of lead (II) iodide precipitated for the reaction.

**Solution**

q = mc∆T

q = mass × specific heat capacity × change in temperature

q = [mPb(NO3)2(aq) + mKI(aq)] × C × (Tf - Ti)

In this case mass = volume (when density is unity).

q = [50 + 30] × 4.184 × (22.2 - 19.6)

= 870.27J

Calculate the moles of species specified, n(PbI2(s)):

From the equation, 1 mole Pb(NO3)2(aq) reacts with excess KI(aq) to produce 1 mol PbI2(s)

Moles Pb(NO3)2(aq) = moles PbI2(s)

n(Pb(NO3)2(aq)) = n(PbI2(s))

n(Pb(NO3)2(aq)) = molarity × volume

n = moles (Pb(NO3)2(aq))

= moles (PbI2(s))

= 0.20 × 50 × 10-3

= 0.010 mol

Calculate the enthalpy of precipitation (heat of precipitation), ΔH, in kJmol-1 of PbI2(s):

ΔH = -q/1000 ÷ n

ΔH = -0.870/1000÷0.010

= - 87 kJmol-1

ΔH is negative because the reaction is exothermic (energy was released causing the temperature to increase).

**Calculating the energy released or absorbed without the specific heat capacity**

The energy released or absorbed during a chemical reaction can be calculated using the stoichiometric coefficients (mole ratio) from the balanced chemical equation and the value of the enthalpy change for the reaction (ΔH):

Energy = n × ΔH

Where:

Energy = energy released or absorbed measured in kJ

n = amount of substance measured in moles

ΔH = enthalpy change for the reaction measured in kJmol-1

**Solved Example 1:**

The synthesis of ammonia gas (NH3(g)) from nitrogen gas (N2(g)) and hydrogen gas (H2(g)) releases 92.4kJmol-1 of heat energy as shown by the balanced chemical equation below:

N2(g) + 3H2(g) → 2NH3(g) ΔH = −92.4kJmol-1

How much energy would be released if we start this reaction with 10 moles of N2(g) and 30 moles of H2(g).

**Solution**

N2(g) + 3H2(g) → 2NH3(g) ΔH = −92.4kJmol-1

This means that:

✓ 92.4 kJ of energy is released for every 1 mole of N2(g) consumed.

✓ 92.4 kJ of energy is released for every 3 moles of H2(g) consumed.

✓ 92.4 kJ of energy is released for every 2 moles of NH3(g) produced.

If 1 mole of N2(g) produced 92.4 kJ of energy, then 10 times as much N2(g) (10 moles of N2(g)) should produce 10 times as much energy:

10N2(g) + 30H2(g) → 20NH3(g) energy = 10 × ΔH

Energy released = 10 × ΔH

=10 mol × 92.4 kJ mol-1

=924kJ

It is useful to consider first how much energy would be released by just 1 mole of H2(g).

We can do this by dividing the entire chemical equation by the stoichiometric coefficient (mole ratio) of H2(g) in the chemical equation, that is, divide every term in the chemical equation (including the value of ΔH) by 3:

Original chemical equation:N2(g) + 3H2(g)→2NH3(g) ΔH = −92.4kJmol-1

Divide by 3: 1/3N2(g) +3/3H2(g) → 2/3NH3(g) ΔH = −92.4/3kJmol-1

1/3N2(g) + H2(g) → 2/3NH3 (g)ΔH = − 30.8kJmol-1

From this new balanced chemical equation we see that when 1 mole of H2(g) is consumed (with excess N2(g)) to produce 2/3 moles of NH3(g), 30.8 kJ of energy will be released.

If 1 mole H2(g) produced 30.8 kJ of energy, then 10 times the amount ofH2( g) will produce 10 times as much energy:

Energy released =10 mol × 30.8kJmol-1 = 308kJ

If we think about what we have done here we have:

✓ divided ΔH value by the stoichiometric coefficient of H2(g) in the balanced chemical equation.

✓ multiplied this new value by the moles of H2(g) actually present. We can generalize this for any balanced chemical equation:

aA + bB → cC + dD …………......................................... ΔH = +kJmol-1

**To calculate the energy absorbed for n moles of reactant A:**

✓ Divide ΔH value by a energy absorbed per mole of A = ΔH/a kJmol-1

✓ Then multiply by **n**, energy absorbed for **n** moles of A = n × ΔH/ a kJ

**To calculate the energy absorbed for n moles of reactant B:**

✓ Divide ΔH value by **b** energy absorbed per mole of B = ΔH/b kJmol-1

✓ then multiply by n, energy absorbed for n moles of B = n × ΔH/bkJ

**To calculate the energy absorbed for n moles of product C:**

✓ Divide ΔH value by **c** energy absorbed per mole of C = ΔH/c kJmol-1

✓ then multiply by **n**, energy absorbed for **n** moles of C = n × ΔH/c kJ

**To calculate the energy absorbed for n moles of product D:**

✓ Divide ΔH value by **d** energy absorbed per mole of D = ΔH/d kJmol-1

✓ then multiply by **n**, energy absorbed for **n** moles of D = n × ΔH/d kJ

The energy released or absorbed during a chemical reaction (energy) can be calculated using the stoichiometric coefficients (mole ratio) from the balanced chemical equation and the value of the enthalpy change for the reaction (ΔH):

Energy = n × ΔH /stoichiometric coefficient

Where;

Energy = energy released or absorbed measured in kilojoules

n = amount of substance measured in moles

ΔH = enthalpy change for the reaction measured in kJmol-1

**Stoichiometric coefficient is that for the particular reactant or product of interest.**

**Calculating Heat Released or Absorbed for a Given Amount of Reactant or Product**

**Solved Example 2:** When sulphur (VI) oxide gas (SO3(g)) is synthesized from sulphur (IV) oxide gas(SO2(g)) and oxygen gas (O2(g)), 198 kJmol-1 of heat is released as shown in the balanced chemical equation below:

**2SO2(g) + O2(g) → 2SO3(g) ΔH = −198 kJ mol-1**

Determine the amount of heat absorbed in kilojoules when 20.0g moles of sulphur (VI) oxide gas decomposes to produce sulphur (IV) oxide gas and oxygen gas.

**Solution:**

The heat absorbed in kJ for the decomposition of 20.0 g, SO3(g) =?kJ

1. Balanced chemical equation for the synthesis of SO3(g):

2SO2(g) + O2(g) → 2SO3(g) ΔH = −198kJmol-1

✓ Reaction for decomposition of SO3(g) is the reverse of the reaction for synthesis of SO3(g) so the sign of the ΔH term is reversed:

ΔH decomposition = −ΔH synthesis

ΔH decomposition = −ΔH synthesis = − (−198) kJmol-1 = +198 kJmol-1

Decomposition reaction: 2SO3(g) → 2SO2(g) + O2(g) ΔH = +198 kJmol-1

✓ Calculate amount in moles of SO3(g):

Molar mass of SO3 = 32 + 16 × 3 = 80 gmol-1

Substituting the values into the above formula, we get;

✓ Calculate energy absorbed by decomposition of **n** moles of SO3(g):

Energy absorbed = n(SO3(g)) × ΔH divided by stoichiometric coefficient of SO3(g)

**Example 3:** A 40.0 mL sample of 0.100 mol/L Ba(NO3)2 is mixed with 20.0 mL of excess NaOH(aq) in calorimeter. During the reaction, there is a rise in temperature by 5.0oC. Calculate the molar enthalpy change for the above reaction.

**Solution**

First, we should find the heat of the reaction. The appropriate equation is q = mC∆T

q = heat = ?, m = mass = 40 ml + 20 ml = 60 ml x 1 g/ml = 60 g (note: we are assuming a density of 1 g/ml for water and we are ignoring the masses contributed by the reagents and the products of the reaction. These are common assumptions to make in this type of problem, especially in Introduction to Chemistry.)

C = specific heat = 4.184 J/g (for water; another assumption), ∆T = change in temperature = 5.00 ˚C

q = (60 )(4.184)(5) = 1255 J. This is the heat (enthalpy) of reaction.

Since they ask for the molar enthalpy, we need to find the moles of Ba(OH)2 formed.

**BOND ENERGY OR BOND ENTHALPY**

Bond energy or bond enthalpy is the energy required per mole of gaseous compound to break a particular bond to produce gaseous fragments at 25°C and 1 atmosphere pressure. It is given in kJ /mole.

> ΔH is positive for breaking bonds.

> Breaking of bonds is an endothermic reaction.

>Bond energies or bond enthalpies are positive.

**Making chemical bonds releases energy (exothermic reaction).**

> Bond formation releases energy.

> ΔH is negative.

> Bond formation is an exothermic reaction.

**ΔH making bonds = -ΔH breaking bonds = -bond energy (= -bond enthalpy)**

**Bond energies, bond enthalpies, can be used to indicate how stable a compound is or how easy it is to break a particular bond:**

>The more energy that is required to break a bond, the more stable the compound will be.

> A larger bond energy implies the bond is harder to break, so the compound will be more stable.

> When comparing two compounds, the compound containing the bond with the lowest bond energy will be the least stable.

✓ A chemical reaction will be endothermic if the energy absorbed to break bonds in the reactant molecules is greater than the energy released when bonds are formed in the product molecules.

✓ A chemical reaction will be exothermic if the energy absorbed to break bonds in the reactant molecules is less than the energy released when bonds are formed in the product molecules.

**CALCULATING THE ENERGY CHANGES IN REACTIONS**

**Bond energy (kJ / mole)**

H-H 436

Cl-Cl 242

H-Cl 431

C-C 346

C=C 612

C-O 358

C-H 413

O=O 498

O-H 464

N=N 946

N-H 391

**Solved Example 1:** The exothermic reaction between hydrogen and chlorine:

H—H + Cl—Cl => 2 H—Cl

Energy in to break each mole of bonds:

1 × H—H = 436 kJ

1 × Cl—Cl =242 kJ

Total energy in = 678 kJ (346 + 242)

Energy out from the two moles of bonds forming:

2 × H—Cl = 2 × 431 = 862 kJ

**Energy in - energy out**

678 kJ - 862 kJ = -**184 kJ (Exothermic)**

So the reaction gives out 184 kJ of energy, overall. ,

**Solved Example 2:** The endothermic decomposition of ammonia:

2NH3 => N2 + 3H2

Energy in to break the two moles of bonds:

6 × N—H(2 × 3N—H) = 6 × 391 = 2346 kJ

Energy out from the four moles of bonds forming:

1 × N = N = 1 ×946 = 946 kJ

3 × H—H = 3 × 436 =1308 kJ

Total energy out = 2254 kJ

**Energy in - Energy out** = 2346 kJ - 2254 kJ = +92 kJ.

So the reaction takes in 92 kJ of energy, overall (Endothermic).

√ is the driving force of the Chemical reaction. It is a state function that relates enthalpy and entropy. A change in Gibbs free energy is represented by ∆G.

∆G = ∆H - T∆S

√ **Entropy** is the measure of the degree of disorderliness of a chemical system. A change in entropy is represented by ∆S.

∆S = SPrdt - SRxt

√ **Enthalpy** is the heat accompanying a chemical reaction. The heat that is evolved or absorbed in the reaction. A change in enthalpy is represented by ∆H.

**∆H = Hprdt - HRxt**

**√ Spontaneous process** is a process which can proceed in a given direction without the need of outside energy.

When:

* ∆G = +ve; the reaction is non spontaneous.
* ∆G = -ve; the reaction is spontaneous.
* ∆H = + ve; the reaction is endothermic.
* ∆H = - ve; the reaction is exothermic.
* ∆S = + ve; the reaction is spontaneous, since there is an increase in entropy.
* ∆S = - ve; a decrease in entropy and it is non spontaneous.

**Example 1:**  Calculate Gibbs free energy change, ∆G given the reaction at s.t.p:

N2(g)  + O2(g)  ==> 2NO(g) ∆H° for N2 and O2 = 0 KJ/mol, ∆H° for NO = 90.37 KJ/mol, ∆S° for N2 = 191.5J/mol.K, O2 = 205 J/mol.K and NO = 210.62J/mol.K.

**Solution**

∆H = Hprdt - HRxt

∆H= 2[90.37] - 0

∆H= 180.74 KJ/mol

∆S = SPrdt - SRxt

∆S = 2[210.62] - [191.5 + 205]

= 421.24 - 396.5

= + 24.74 J/mol.K

= + 0.02474KJ/mol

**To find the Gibb's free energy, we use the formula below:**

∆G = ∆H - T∆S

At s.t.p, the absolute temperature = 273K.

∆G= 180.74 - (273 × 0.02474)

= 180.74 - 6.75402

∆G = + 173.98598

**Because this value is positive, we know the forward reaction is non spontaneous (not feasible).**

**Example 2:** How much heat is released by burning 27.5g of methane?

CH4(g) + 2O2(g) =>CO2(g) + 2H2O(l) ∆H= -890.4KJ

**Solution**

CH4(g) + 2O2(g) => CO2(g) + 2H2O(l)

Find the molar mass of CH4 :12 × 1 + 1 × 4 = 16g

16g CH4 = - 890.4KJ

27.5g CH4 = ?

Cross multiplying;

27.5 × - 890.4/16 = - 1530.375KJ

**The reaction is exothermic since ∆H is negative.**

**Example 3:**

C2H4(g) + 3O2(g)=> 2CO2(g) + 2H2O(l)

**∆Hf°(KJ/mol) +52.30 - 393.5 - 241.8**

Calculate the enthalpy change of the above reaction (∆Hrxn)

**Solution**

∆Hrxn = sum of Hprdt - sum of Hrxt

= [2(-393.5) + 2(-241.8)] - [52.3 + 3(0)]

= - 1322.9KJ/mol.

**The reaction is exothermic because ∆H = - ve.**

### Note: The coefficients of the reactants and products are taken into consideration during the calculation.

**Example 4:** What is the enthalpy change for the reaction below?

CH4(g) + Cl2(g) => CH3Cl(g) + HCl(g)

If the enthalpy of CH4, CH3Cl and HCl are -74.9 kJmol-1, -83.7 kJmol-1 and -92.0 kJmol-1 respectively.

**Solution**

∆Hrxn = sum of Hprdt - sum of Hrxt

∆Hrxn = - 83.7 - 92.0 - (- 74.9)

∆Hrxn = - 83.7-92.0 + 74.9

∆Hrxn = - 100.8 kJ

**Example 5:** How much energy is needed to convert 180 grams of ice at 0aC to liquid water at the same temperature? The heat of fusion for water is 6.0kJ/mol.

**Solution**

Molar mass of H2O = 1 mass of H𝑠𝑠 𝑜𝑓 𝑤-1

We now can calculate the energy needed for the conversion:

1 mole of water = 6.0 kJ/mol

10 moles of water = ?

Cross multiplying, we get;

10 moles ×oles J/mol / 1 mole = 60kJ

**TAKE AWAY**

1. A student of Ray-Field International Secondary School, Akwa Ibom State performed an experiment to determine the enthalpy change in the reaction of magnesium with dilute HCl. He added 5g of magnesium ribbon to excess dilute hydrochloric in a glass Calorimeter. The temperature of the reaction system rose from 25°C to 45°C at the end of the reaction. Assuming that the specific heat capacity of HCl is 4.2 Jg-1K-1. Calculate the:

I) Enthalpy change of solution of magnesium ribbon. **Ans. 6153Jg-1K-1.**

II) Molar enthalpy change of solution of magnesium ribbon in KJmol-1.**Ans. 29.5817 KJmol-1.**

III) State 3 sources of errors in the student's experiment. **Ans.:**

**a. Improper mixing of the reactants.**

**b. Placing the thermometer bulb slowly on dissolving particles.**

**c. Inaccurate thermometer reading.**

**d. Heat loss since the experiment was not carried out in a closed system.**

2. How much energy is needed to convert 90 grams of ice at -12sed system.C at the end of the heat of fusion for water is 6.0kJ/mol. **Ans. 30 kJ**

3. In a laboratory experiment, 1.16g of an organic liquid fuel was completely burned in oxygen. The heat formed during this combustion raised the temperature of 100 g of water from 295.3 K to 357.8 K. Calculate the standard enthalpy of combustion of the fuel. Its relative molecular mass is 58.

**Ans. -1306 KJmol-1**

CHAPTER 13: RADIOACTIVITY

INTRODUCTION

is the process in which a nucleus spontaneously disintegrates, giving of radiation. The radiation consists of one or more of the following, depending on the nucleus: electrons, nuclear particles (such as neutrons), smaller nuclei (usually helium - 4 nuclei), and electromagnetic radiation.

Rutherford used a radioactive element as a source of alpha particles and allowed these particles to collide with nitrogen nuclei. He discovered that protons are ejected in the process. The equation for the nuclear reaction is

+–> +

The experiments were repeated on other light nuclei, most of which were transmuted to other elements with the ejection of a proton. These experiments yielded two significant results. First, they strengthened the view that all nuclei contain protons. Second, they showed for the first time that it is possible to change one element into another under laboratory control.

When beryllium is bombarded with alpha particles, a penetrating radiation is given off that is not deflected by electric or magnetic fields. Therefore, the radiation does not consist of charged particles. The radiation from beryllium consists of neutral particles, each with a mass approximately that of a proton. The particles are called neutrons. The reaction that resulted in the discovery of the neutron is

+ –>+

**BALANCING OF NUCLEAR EQUATIONs**

1. Balance the following nuclear reaction:

Use to represent a neutron andto represent an electron.

**Solution**

9 + 4 = 12 + 2 =13 => mass number

4 + 2 = 6 => atomic number

Since the sum of atomic numbers of the reacting nuclei is the same as the atomic number on the atomic number of the product and a difference of one in the mass number, use (neutron) to balance the nuclear equation.

**Example 2:**

**Solution:**

To balance the mass number:

Here you ask yourself what number would I add to 4 to give 212 (you add 208) and again to balance the atomic number, what number will you to add to 2 to give 84 (you add 82). Then, use the periodic table to find the element with the atomic number of 82 which is Pb (lead).

**Example 3:**

**Solution:**

Since the sum of the atomic number of the reacting nuclei is the same as the atomic number of the product, useto balance the equation.

253 + 4 = 1+256 = 257 => mass number

99 + 2 = 0 + 101 = 101 => atomic number

The total mass of reactants is 257 and total atomic number is 101.To complete the balancing; add mass of 256 to 1 and 101 to 0. Then, use the periodic table to find the element with the atomic number of 101 which is Md (Molydenum).

**Example 4**: Balance the nuclear equation below:

**Solution:**

The mass of reactant is same as that of product, but the only difference is in atomic number by decrease of 1. Here, we have electron capture because there is no charge in the mass number but the atomic number decreases by 1().

**TYPES OF RADIATION**

There are 3 types of radioactive decay:

***1. Alpha decay (α)***: When an alpha particle is injected into an unstable nucleus (). An alpha particle has the same composition as helium nucleus.

E.g Balance:

**Solution:**

There are 238 mass number on the right; we have already 4 mass number so we need 234 mass number on the left. For the atomic number (proton number), there are 92 protons on the right, we have already 2 protons. So we need 90 protons on the left to accurately balance the equation. Use your periodic table to check the element with 90 protons (thorium).

**2. Beta decay *():*** Here, an electron is ejected from the nucleus. An electron is represented as ()

E.g balance the equation: =>  + ?

There are 234 mass number on the left but we have already zero mass on the right, so we need 234 on the right to balance the mass number. For the atomic number, we have 90 protons on the left and on the right we have a negative 1 (-1), so we must add 91 protons such that -1 plus 91 will give 90 (same as on the left). When you look at the periodic table, protactium is number 91. Hence, =>+

**Note:** = +

**3. A gamma ray, or g ray**, refers to a high-energy photon. The emission of gamma rays is one way a nucleus with excess energy (in an excited nuclear state) can relax to its ground state. Gamma rays have no mass number and no protons (no charge).

E.g balance => + ?

**Solution**

=> +

Excited state Ground state

**HOW TO WORK OUT THE ENERGY RELEASED IN NUCLEAR REACTIONS**

**Some common terms**

1. **Nuclear Binding Energy:** Nuclear Binding Energy of a nucleus is the energy needed to break a nucleus into its individual protons and neutrons.

2. **Mass defect:** Mass defect is the difference between the calculated mass of the unbound system and the experimentally measured mass of the nucleus.

3. **Nuclear fusion:** Nuclear fusion is a nuclear reaction in which light nuclei combine to give a stable, heavier nucleus plus possibly several neutrons, and energy is released.

4. **Nuclear fission:** Nuclear fission is a nuclear reaction in which a heavy nucleus splits into lighter nuclei and energy is released.

**ENERGY RELEASED IN NUCLEAR REACTIONS**

The energy released during nuclear reactions is given by the formula below:

E = mc2

Here **C** is the speed of light, 3.00 × 108 m/s, **E** is the energy released and **m** is the mass defect of the nucleus.

**Example 1:**

**For Nuclear fusion:** In this reaction, two isotopes of hydrogen i.e deuterium and tritium react together to form helium and neutron. By calculating the total mass of the reactants and total mass of the products, we can get the mass defect. We subtract the mass of products from the mass of reactants (mass defect). We can use the equation E = mc2 to calculate the energy released in the nuclear fusion reaction.

+  => +

**2.013553 3.016049** => **4.001506 1.008665**

5.029602 u => 5.010171 u

**Note:** 2.013553 + 3.016049 gives 5.029602 u while 4.001506 + 1.008665 gives 5.010171 u.

One unified atomic mass unit = 1.66 × 10-27Kg. Mass defect = total particle (mass of reactants minus total particle mass of products).

Mass defect = (5.029602 - 5.010171) u

Mass defect = 0.019431 u

Mass defect = 0.019431 × 1.66×10-27

Mass defect = 3.23 × 10-29 Kg

To calculate the energy released, use E = mc2

E = 3.23 × 10-29 × (3.00 × 108)2

E = 2.91 × 10-12J (One helium nucleus)

For one mole of 4He:

2.91 × 10-12 × 6.02 × 1023 = 1.75 × 1012 KJmol-1

**Example 2:** Calculate energy released during nuclear fusion of.

+ => + + 3

[Mass of U = 235.0439, = 1.008665, Ba = 140.9144, Kr = 91.92616]

**Solution**

Mass defect = Mass of reactants - mass of products

+ => ++ 3

**235.0439 + 1.008665 => 140.9144 + 91.92616 + 3.02599**

**236.0526 u =>235.8666 u**

Mass defect = 236.0526 - 235.8666 = 0.186u

**Recall: One unified atomic mass unit (u) = 1.66 × 10-27Kg.**

0.186 × 1.66 × 10-27 = 3.0876 × 10-28Kg

E = mc2

3.0876 × 10-28Kg × (3.00×108)2

3.0876×10-28  × 9.0 × 1016

E = 2.78 × 10-11J (**Energy released for one atom of 235U**)

Then, the energy for one mole of 235U:

2.78 × 10-11 × 6.02 × 1023 = 1.67 × 1010KJmol-1

**HALF - LIFE ()**

The half life of a radioactive nucleus is the time it takes for one half of the nuclei in a sample to decay. The half life is independent of the amount of sample.

**Note:** In some cases, you easily find the half life of a radioactive sample by directly observing how long it takes for one half of the sample to decay. For instance, you find that 1.00g of iodine -131, an isotope used in treating thyroid cancer, decays to 0.5g in 8.07 days. Therefore, its half life is 8.07 days since one-half of 1.00g is 0.5g.

**Some useful equations:**

Where K = decay constant, t1/2 = half-life.

1. A = Ao • =)t/h

Where A = final amount, Ao = initial amount, 1/2 = Split factor (After a half life, one gram becomes 1/2 gram), t = time, h = half life.

1. AF = Aoe-kt

Where AF = Final amount, Ao = initial amount, t =time of decay, k = Decay constant.

Where T = half life, t = time, N1 = original mass, N2 = mass remaining after decay.

**Example 1**: The decay constant for the beta decay of is 1.0 × 10-13/s. What is the half life of this isotope in years?

**Solution**

Then, you can convert this half life (6.93 × 1012 ) to years:

60 seconds = 1 minute

60 minutes = 1 hour

24 hours = 1 day

365 days = 1 year

Therefore, 60 seconds = 1 minute

6.93 × 1012 seconds = ?

Cross multiplying we get;

1 minute × 6.93 × 1012 s/ 60 seconds = 115000000000 minutes

Convert the above minutes to hours:

60 minutes = 1 hour

115000000000minutes = ?

Cross multiplying we get;

115000000000 minutes × 1 hour /60 = 1916666666.7 hours

Convert hours to days:

24 hours = 1 day

1916666666.7hours = ?

Cross multiplying we get;

1916666666.7 hours × 1 day/24 hours = 79861111.113 days

Convert days to years;

365 days = 1 year

79861111.113days = ?

Cross multiplying we get;

79861111.113days × 1 year / 365 days = 224328.9638 years

2.2 × 105 years

**Example 2:** An isotope has a half-life of 6 hours. What percentage will be left after 24 hours?

**Solution:** Using timeline

| Half life | 0 | 1 | 2 | 3 | 4 |
| --- | --- | --- | --- | --- | --- |
| Time (Hours) | 0 | 6 | 12 | 18 | 24 |
| Percentage | 100 | 50 | 25 | 12.5 | 6.25 |

That is 6.25 % will be left after 24hours. At time zero (0), the half life is zero (0) and the percentage is 100. Each half life is 6 hours, so we have 6, 12, 18 and 24. Half of 100 = 50%, half of 50% = 25%, half of 25% = 12.5% and half of 12.5% = 6.25%. Hence, after 24 hours 6.25% will be left.

**Example 3:** The half-life of Pd-100 is 4 days. After 12 days, a sample has been reduced to a mass of 4g. What is the starting mass?

**Solution:**

| (Half) | Time (days) | Amount (g) |
| --- | --- | --- |
| 0 | 0 | 32 |
| 1 | 4 | 16 |
| 2 | 8 | 8 |
| 3 | 12 | 4 |

The starting half life is zero (0), time in days in zero (0) and the amount is unknown. We know that the half life is 4 days; we have to go down to 12 days by simply adding 4 till we get 12 days (4 + 4 + 4 = 12 days). We know that we are left with 4g after 12 days, so that will mean 3 half life. We go backward from the 4g and multiply it by 2(8g), we multiply 8 by 2(16g) and finally multiply 16 by 2 to have 32g.

**Example 4:** How long will it take for a 64g sample of Rn – 222 that has a half-life of 3.8 days to decay to 8g?

**Solution:**

|  | Time (days) | Amount(g) |
| --- | --- | --- |
| **0** | **0** | **64 ÷4o** |
| **1** | **3.8 + 3.8** | **32 ÷23** |
| **2** | **7.6 + 3.8** | **16 ÷6** |
| **3** | **11.4** | **8** |

At zero half life, the time = 0 and the amount 64g. We know that1 half-life is 3.8 days. By definition, half life means time taken for half of the product to decay, so half of 64 is 32(divide 64 by 2 = 32g). For second half-life, we are going to add 3.8(3.8 + 3.8 = 7.6) and half of 32g = 16 g. For the third half-life, we are going to add another 3.8(7.6 + 3.8 = 11.4) and the half of 16 g = 8g.

Hence, it will take 11.4 days for a 64g sample of Rn- 222 to decay to 8g.

**Example 5:**

Iodine -131 has a half-life of 8days. If there are 200 g of this sample, how much of iodine -131 remains after 32 days?

**Solution:**

AF=Aoe-kt

Where AF = Final amount, Ao = initial amount, t = time of decay, k = Decay constant

AF = 200e[-0.08664 ter.6

= 12.50g

OR

8days 8days 8days 8days

200g 100g 50g 25g 12.5g

Hence after 32days (8 + 8 + 8 + 8 days), 12.5g will remain.

**Example 6:**

Sodium- 24 has a half-life of 15hours. If there are 800g of Na-24 initially, how long will it take for 750 g of Na-24 to decay?

**Solution:**

15hrs 15hrs 15hrs 15hrs

800g 400g 200g 100g 50g

It takes four half-lives for the sample to decay I.e 15 + 15 + 15 + 15 = 60 hrs. It is going to take 60 hours for 800g to decay to 50g.

OR

To find the time, use In(AF/Ao) = -kt

AF = 800g - 750g = 50g

In(50/800) = 0.04621 0)g

t = 59.9998 hours ~ 60 hours

**Example 7:** The half-life of oxygen- 15 is 2 minutes. What fraction of a sample of oxygen – 15 will remain after 5 half-lives?

**Solution**

2 mins 2mins 2mins 2mins 2mins

100% 50% 25% 12.5% 6.25% 3.125%

Hence, after 5 half-lives (2 + 2 + 2 + 2 + 2 = 10 minutes), 3.125% will remain.

Convert the % to decimals;

3.125% = 0.03125 = 1/32

**Example 8:** It takes 35 days for a 512g sample of element X to decay to a final amount of 4g. What is the half life of element X?

**Solution**

In(AF/Ao) = -kt

In(4/512) = -k (35)

K = ln(4/512)/ - 35

K = 0.13863

t1/2 = 0.693/K

t1/2 = 0.693/0.13863 = 4.999975

t1/2 = 5 days

**Example 9:** The half life of uranium is 24 days. Calculate the mass remaining unchanged of 0.64g of the substance after

I. 24 days

II. 48 days

III. 72 days

IV. 96 days

V. 120 days

**Solution**

I. After 24 days; 1/2 × 0.64 = 0.32g remains

II. After 24 days;1/2 × 0.32 = 0.16 g remains

III. After 24 days;1/2 × 0.16 = 0.08 g remains

IV. After 24 days;1/2 × 0.08 = 0.04 g remains

V. After 24 days;1/2 × 0.04 = 0.02 g remains

OR

Where T = half life, t = time, N1= original mass, N2 = mass remaining after decay.

I. t = 24 days, T = 24 days

N2 = N1/2t/T = 0.64/224/24

N2 = 0.64/21 = 0.32g

II. t = 48 days, T = 24days

N2 = N1/2t/T = 0.64/248/24

N2 = 0.64/22 = 0.64/4 = 0.16g

III. t = 72days,T = 24days

N2 = N1/2t/T = 0.64/272/24

N2 = 0.64/23 = 0.64/8 = 0.08g

IV. t = 96 days, T = 24days

N2 = N1/2t/T = 0.64/296/24

N2 = 0.64/24 = 0.64/16 = 0.04g

V. t = 120 days, T = 24days

N2 = N1/2t/T = 0.64/2120/24

N2 = 0.64/25 = 0.64/32 = 0.02g

**TAKE AWAY**

1. Calculate the average binding energy per mole of a U-235 isotope. Show your answer in kJ/mole. **Ans. E = 2.7843 x 10-10 J**

2. A radioactive element has a decay constant of 0.077S-1. Calculate its half life.

**Ans. = 9 seconds**

3. What is the decay constant of a radioactive element whose half life is 3 seconds? **Ans. 0.231 sec-1.**

4. The half life of a radioactive element is 5 seconds. Calculate its decay constant. **Ans. 0.1386sec-1.**

5. Uranium of atomic number 92 and mass number 238 emits an alpha particle from its nucleus. The new nucleus formed respectively, atomic number ----- and mass number -----. **Ans. 90 and 234 respectively.**

6. The half life of a radioactive substance is 5 hours. If 5g of the substance is left after 30 hours, determine the original mass of the substance. **Ans. 80g**